ELEMENTARY GENERAL SCIENCE

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PREFACE

WITHIN the past few years distinct advances have been made in the methods of teaching science. It is becoming more and more widely recognised that the only scientific knowledge worth having is that gained by individual observation and experience; and public examining bodies are one by one bringing themselves in line with this sound educational principle by causing their examinations to be more a test of familiarity with experimental methods and results than of information acquired by reading. The new regulations for the Matriculation Examination of the London University, making an experimental knowledge of the fundamental principles of physical and chemical science compulsory for every candidate, represent an advance in this direction. Moreover, provision for such an introductory course in science is made in the Junior Local Examinations of the Universities of Oxford and Cambridge; and pupil teachers in elementary schools are now encouraged by the Education Department to take up a course of elementary science having nearly the same scope as that just introduced into the Matriculation Examination of the University of London.

The lessons in this volume follow the Matriculation Syllabus of "General Elementary Science." It is distinctly stated in a note prefixed to the syllabus that "In particular, the subjects of the present syllabus will be treated wherever possible from an experimental point of view. Candidates will be expected to have performed or witnessed simple experiments in illustration of the subjects mentioned in this syllabus." With these requirements in mind, we have made experimental work a distinctive feature of the book. A large number of simple and practicable experiments, which, with a few exceptions, can be easily performed by the individual student, are described; and our object has been to make the volume a combined class-book and practical exercise book on the rudiments of physical and chemical science.

Many of the experiments follow the course of instruction in elementary science contained in the physics and chemistry syllabus prepared by a Committee of the Incorporated Association of Head Masters. The results are indicated in each case, hence the experiments may be regarded in a way as solved examples of the general laws and phenomena of physical science, but the examples can only be properly understood by actually following the various stages which lead to their solution. In a few sections of the book it has been necessary to illustrate by numerical examples the principles involved; and to exercise the student in solving simple problems of this kind numerous questions have been added to the chapters. For a large number of these, and more especially for those appended to Chapters III—VII, we are indebted to Mr. Walter J. Wood, B.A., who has collected many from actual examination papers, and has supplied the answers.

Throughout the preparation of the following pages we have found the assistance and advice of Professor R. A. Gregory invaluable. Teachers who know Professor Gregory's *Exercise Book of Elementary Practical Physics* will recognise in the earlier chapters of the present volume several experiments described in that book. The opportunity of expressing thanks for permission to introduce these is gladly taken; and we also desire to express our obligations to the Publishers for permitting free use of suitable illustrations from several of their elementary books in the various branches of science with which this volume is concerned.

It is worth remark that although the book is, in a broad sense, a joint production, one of us (A. T. S.) is more directly responsible for Chapters I—XIII, and the other (L. M. J.) for Chapters XIV—XXIII.

A. T. SIMMONS.
L. M. JONES.

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ELEMENTARY GENERAL SCIENCE

CHAPTER I

PHYSICAL PROPERTIES AND STATES OF MATTER

What is meant by "Matter."—Our earliest knowledge of the world teaches us that on every side we have what we familiarly speak of as *things* of all kinds. We become aware of the existence of these things in different ways. Some we feel, some we smell, some we see, some we taste, while others again make their existence known to us by the sounds we hear. On a windy day at the sea-side, standing on the beach, we feel the ground under our feet; we smell, it may be, the tar on a neighbouring boat or the seaweed on the shingle; we see a distant ship at sea or the clouds hurrying across the sky; we taste the salt in the air; and we hear the never-ceasing roar of the waves as they break in at our feet. All these things, about which we get to know by our senses, are called *material* things; they are forms of *matter*. We must think of **matter**, then, as meaning **all things which exist in or out of our world, which we can become aware of by the help of our senses.**

Different kinds of Matter.—Of course the number of different kinds of things is innumerable, but yet they can all be arranged in three classes, according to certain of the properties they possess and which we shall immediately have to study. The classes are (1) Solid things or solids; (2) Liquid things or liquids; (3) Gaseous things or gases. Sometimes the last two are made into one class and called fluids.

What is meant by "Properties."—We shall have occasion to use the word *properties* so often that it will be well to clearly understand what meaning the word conveys, and this can be best accomplished by one or two examples. We say a strawberry is sweet, or a strawberry has the property of sweetness; the paper of the book is white, or the paper possesses the property of whiteness; the sun is bright, or the sun is noted for the property of brightness. Evidently, then, "*properties are certain effects caused by the things which are said to possess them.*"

Properties possessed by all kinds of Matter.—There are certain properties possessed in common by all kinds of matter; these are said to be *general properties*.

1. *Matter must occupy a certain space*, or possesses *extension*; the larger it is the larger the space occupied by it.

2. It will further be obvious to every one that *two material things cannot occupy the same space at the same time*. This property is expressed by saying that matter is *impenetrable*. To be quite accurate, we have to know something about the way in which matter is built up. We shall see more fully later that it is generally supposed that every kind of matter is built up of small constituent parts, which are incapable of division, and that it is really only these indivisible parts which are impenetrable.

3. *Matter offers resistance.*—We become aware of this, in the case of solids, if we knock ourselves against the wall or the table; if we swim or wade in water we know the same thing is true of water, and so we find it to be of all liquids; if we attempt to run with a screen in front of us we become conscious of the resistance offered by the air to our onward progress, and from this argue that gases, too, offer resistance.

4. *Matter has weight.*—Without knowing the full significance of the expression *weight*, we shall have a sufficiently clear idea of what is meant by this property from its familiar use in everyday conversation. By lifting a solid we become conscious of its possession of this property; if we lift an empty bottle and then when it is full of any liquid, we shall find it is lighter in the first instance, or, as we say, the liquid has weight. By the exercise of sufficient care, in just the same manner it can be shown that gases have weight.

5. *Matter transfers motion to other things when it strikes against them.*—If we throw a stick at a cocoa-nut at a fair, or send a jet of water at a ball, or blow at a piece of paper, another of the general properties of matter can be demonstrated, namely, the power of giving motion to other things by striking against them.

All these general properties can be brought together in a definition—thus :—"**Matter occupies space, offers resistance, possesses weight, and transfers motion to other things when it strikes against them.**"

Other Properties of Matter.—Matter possesses other characters or properties which it will be useful for us to study. Though these, too, are general properties, it is possible to form a good elementary notion of matter without taking them into account, and it must be remembered that these properties cannot all be applied to every kind of matter. We shall consider (1) Divisibility, (2) Porosity, (3) Compressibility, (4) Elasticity, (5) Inertia.

Divisibility.—Imagine some material body before you on a table. You know that with suitable means you can divide it into parts by cutting, that each of the halves can be again divided, and that the bisection can be continued as long as the knife is sufficiently fine and sharp to be able to cut the substance. Evidently, if you could only get sharper and sharper knives, and keener and keener eyes, this process of division could be carried on for a very long time. This property is what is understood by *divisibility*.

Could this division go on for ever? There are reasons for believing that it could not. You could not go on dividing matter indefinitely; by and by extremely minute and indivisible parts would be reached, called *atoms*. It must be at once understood that atoms have never been seen. We can only imagine what would be the end of our process of division. Our strongest microscopes bring us nowhere near the possibility of seeing an atom.

Porosity.—We are all in the habit of associating this property with certain familiar forms of matter. The sponge we use in the bath has holes through it, or is, as we say, *porous*. A piece of blotting paper is another common example of an obviously porous material; the substances used in filters must also evidently be porous, or else the water would not percolate through them. Porosity refers to the possession of these *interstices* or *pores*. In some cases, though we cannot see these pores with the naked eye, we easily perceive them with a microscope. The pores have often been shown to exist, even where it is difficult to imagine their existence, by forcing water through them. Thus Francis Bacon, in 1640, forced water through a very carefully closed sphere made of lead.

EXPT. 1.—Filter some muddy water through a filter paper placed in a funnel (Fig. 1). The particles of water are small enough to pass through the pores of the paper, but the solid particles of mud are too large, hence they are left behind.

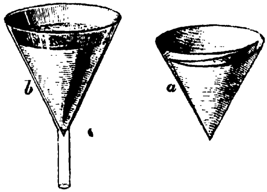

FIG. 1.—Funnel and Filter-paper.

EXPT. 2.—Procure a piece of chamois leather; make it into a bag, and pour some mercury into it. Increase the pressure on the mercury by twisting the leather. The mercury is forced through the pores. This is a common way of filtering mercury.

EXPT. 3.—Half fill a barometer tube with water; then gently add alcohol until the tube is nearly full. Make a mark on the tube at the level with the top of the liquid column, and afterwards shake the tube so as to mix the water and alcohol well together. Observe that the volume of the mixture has diminished, the reason being that some of each liquid has filled up pores between the particles of the other.

Compressibility.—This property follows as a natural consequence of that just described. If pores exist between the indivisible small particles of which matter is built up, it ought to be possible, by the adoption of suitable means, to make these particles go closer together. This is found to be the case. By pressing upon the body from outside it can be made smaller.

This is very well known to be the case in gases, which can actually be made to become successively one-half, one-quarter, one-eighth, and so on up to at least one-hundredth of their original size.

The same thing holds true in the case of solids, though to a much smaller extent. A familiar example of the compression of solids is seen when a druggist presses a cork between two pieces of iron in order to make it fit a bottle for which it was previously too large. But, generally, in the case of solids the pressure exerted has to be very great to bring about even a small compression.

It was believed for a long time that liquids could not be compressed, but it is now known that they can be very slightly reduced in volume, that is, the particles can be forced nearer together.

Thus we learn that *compressibility is not only a consequence of porosity but actually a proof of its existence.*

Elasticity.—Imagine a gas to have been made to assume one-half its size by compressing it. What would happen if the pressure, which is the cause of the diminution, were suddenly removed? The gas would resume its original size or volume, and it would, so far as appearances are concerned, seem to have undergone no change. The gas is said to be perfectly *elastic* and the property which enabled it to go back to its original state is called *elasticity*. Similar results follow with liquids; they also are perfectly elastic.

Some differences arise when solids come to be examined. Though the property can be developed in solids in at least four ways—by *pressure*, by *pulling*, by *bending*, and by *twisting*—we need only in this connection consider the first, as it is the elasticity which is developed by pressure which is most marked in all forms of matter. Ivory, marble, and glass are examples of elastic solids; while putty, clays, fats, and even lead are instances of solids with scarcely any elasticity. In a scientific sense, glass is more perfectly elastic than india-rubber, because it returns to its original shape after it has been forced out of that shape, whereas india-rubber does not exactly return to its original shape.

A solid will only resume its former dimensions when the pressure is removed, provided that the pressure is within a certain limit. If the pressure be more than this minimum amount, or if it exceeds the *limit of elasticity*, as it is called, the solid will not return to the initial size; it will undergo a permanent change. As the student will see later from Expt. 10, this limit of elasticity is only exceeded in the case of india-rubber when the pressure applied is very great.

> Expt. 4.—Procure a slab of polished marble or some similar material and smear it with oil. Drop a billiard ball or a large glass marble from a considerable height on to the slab. Catch it as it rebounds. Notice that a blot of oil is found where the ball came into contact with the slab. Compare the size of the blot with the spot which is formed when the marble is placed in contact with the slab.

Evidently the ball underwent a compression as the result of collision with the slab, and, by virtue of its elasticity, it regained its original size, causing the rebound.

Inertia.—This property will be considered more fully in a later chapter, and we will let it suffice to say here that it is entirely a negative property; it may be expressed in a general way by the statement that inanimate bodies are incapable by themselves of changing their state of rest or motion.

States of Matter.—Solids, Liquids, Gases.—The fact that there are three kinds of material things is well known to every one, and has been stated already. We must now add to this another idea, viz., that **the same matter can exist in three different states.**

Expt. 5.—Procure a lump of ice and notice that it has a particular shape of its own, which as long as the day is sufficiently cold, remains fixed.

Expt. 6.—With a sharp brad-awl or the point of a knife break it up into pieces, and put a convenient quantity of them into a beaker. Place the beaker in a warm room, or apply heat from a laboratory burner or spirit lamp. The ice disappears, and its place is taken by what we call water. Notice the characters of the water. It has no definite shape, for by tilting the beaker the water can be made to flow about.

Expt. 7.—Replace the beaker over the burner and go on warming it. Soon the water boils, and is converted into vapour, which spreads itself throughout the air in the room, and seems to disappear. The vapour can only be made visible by blowing cold air at it, when it becomes white and visible; but is really no longer vapour, but has condensed into small drops of water.

Here the same form of matter has been made to assume three states; in other words, ice, water, and steam are the same form of matter in the solid, liquid, and gaseous state respectively.

The change from one state to another may be sudden or gradual.—The circumstances attending the change from the solid to the liquid, or from the liquid to the gaseous state, are not always the same as we have seen in the case of water. When solid iodine is heated, it appears to suddenly pass from the condition of a solid to that of a gas. Camphor is another instance of this sudden transition from solid to vapour. When, on the other hand, sealing-wax is heated, it very gradually passes into the liquid condition, and may be obtained in a kind of transition stage—neither true solid nor true liquid.

> EXPT. 8.—Warm a Florence flask by twirling it between the finger and thumb above the flame of a laboratory burner. When it is too warm to bear the finger upon the bottom, introduce a crystal of iodine, and notice it is at once converted into a beautiful violet vapour.
>
> EXPT. 9.—Warm a lump of sealing-wax or bicycle cement in an iron spoon, and notice the gradual conversion into a liquid.

There is no hard and fast line between the three conditions of matter we have mentioned. The reader will find as his knowledge extends that intermediate states are known between those we have mentioned, but for the present it will be best to confine the attention to this simple division, reserving for future work a study of these gradations. We shall now consider the distinctive properties of solids, liquids, and gases.

DISTINCTIVE CHARACTERS OF SOLIDS.

A solid body does not readily alter its size or shape. It will keep its own volume and the same form unless subjected to a considerable force.—Another way of expressing the facts contained in the definition is to say that *solids possess rigidity*. Hard solids are more rigid than soft ones. Liquids, on the other hand, possess no rigidity, but the opposite property of flowing or *fluidity*. Liquid particles slide over one another so easily that the surface of a liquid at rest is always horizontal. Fine sand may be made to flow, but the particles, however, do not move freely over one another, so the surface is left uneven, which fact provides us with a clear means of distinguishing between a solid and a liquid.

Solids possess Elasticity.—It has been already shown by Expt. 4 that elasticity can be called into play in solids by pressure, and we have now to prove that it is also called into play by pulling, bending, or twisting.

It would take us farther into the subject of Physics than we can afford space to describe how the elasticity has been measured in these cases, but that there is an alteration of form when solids are treated in these ways can be easily shown.

EXPT. 10. Fix one end of a piece of indiarubber cord, or tubing, about two feet long, to a support. Stick two pins through this cord about 18 inches apart. Tie the lower end of the cord into a loop, and then hang a weight by means of a hook from it. Measure the distance between the pins before and after putting on the weight. Repeat the experiment with different weights. You will notice that the cord can be stretched or elongated, because the weights exert a pulling force upon it. If a long metal wire is used instead of the cord, the stretching can be measured in a similar way, but it is much less; for instance, a brass wire $\frac{1}{25}$ inch in diameter and eleven feet long elongates about $\frac{3}{25}$ inch when a weight of 28 lbs. is pulling it.

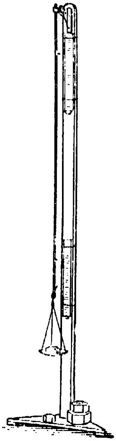

FIG. 2.—Experiment to illustrate the longitudinal stretching of an india-rubber cord.

EXPT. 11. Procure a flexible wooden lath, and fix it horizontally by clamping one end of it firmly. To the other end attach a pin by means of a little wax. Place a rule vertically near the pin, as in Fig. 3. A weight should then be hung from the free end of the lath, and the amount of bending observed. Keeping the same weight, clamp the lath so that only half the previous length can be bent, and again notice the amount of bending. Try also with other lengths.

EXPT. 12. Suspend a wire with a weight at its lower end, and under it a circle divided into degrees in a manner similar

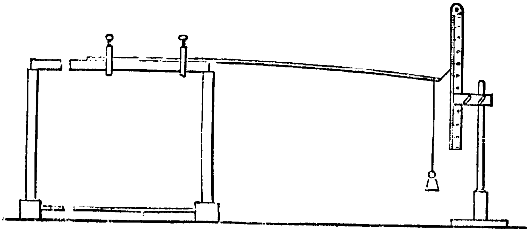

FIG. 3.—Measurement of the Bending of a Lath.

to that shown in Fig. 4. Notice the position at which the weight comes to rest; then twist the weight through a certain angle, and let it go. The weight untwists back to the starting point and beyond it, and then spins in the way it was twisted, and goes on oscillating in this manner until it comes to rest. Observe how long the weight takes to make ten or fifteen complete spins. Repeat the observation with wires of different length and diameter, and made of different metals. The time of an oscillation depends upon the tendency to untwist, so that the experiment shows that this depends, among other things, upon the length, diameter, and nature of a wire.

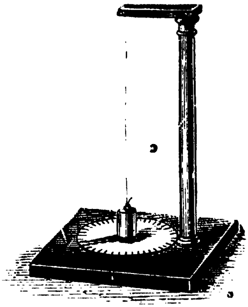

FIG. 4.—Elasticity of Torsion.

Solids possess Tenacity, Ductility and Hardness.—The force required to tear asunder the particles of a body varies with different materials, in other words, some substances are more tenacious than others.

1. *Tenacity is measured by ascertaining what weight is necessary to break solids when in the form of wires.*

EXPT. 13. Suspend a balance-pan from the lower end of a thin copper wire attached to a beam. Add weights to the pan until the wire breaks. The force required to break the wire is the joint weight of the balance-pan and the weights in it. Repeat the experiment with wires of the same diameter but made of different material.

In making the measurement of tenacity, the area of the cross section of the wire must first be carefully estimated. By the cross section is meant the area of the end of the wire when it is carefully filed to be at right angles to the length. This cross section is estimated by measuring the diameter of the wire and calculating.[1]

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[1] Area of cross section $= \frac{22}{7} \times \frac{d^2}{4}$

It is found that a wire of twice the cross sectional area of another will be just twice as tenacious. Evidently, then, if we wish to compare the tenacity of two wires of different materials, it will make the experiment much simpler if wires of the same cross section are selected. Cast steel is the most tenacious of all metals, being about twice as much so as copper and forty times as tenacious as lead. But the tenacity of steel itself is exceeded by that of unspun silk, while single fibres of cotton can support millions of times their own weight without breaking.

2. *Ductility is the property by virtue of which solids can be made into wires.* A ductile material is thus one which can be drawn out. The change of form in this case is produced by pulling. *Malleability* is a similar property to ductility, but the change of form is brought about by the application of pressure; gold, copper, and lead, for instance, can be beaten out into thin plates, and are therefore malleable substances. Lead is an example of a malleable material which is not ductile—it can be beaten out but cannot be drawn into wires.

Platinum is the most ductile and gold the most malleable metal known. Platinum has been drawn out into wire so fine that a mile of it weighs only one and a quarter grains. Gold has been beaten into plates so thin that it would require three hundred thousand of them placed one above the other to make a layer an inch thick.

3. *Hardness is the property by virtue of which solids offer resistance to being scratched or worn by others.* This is a property of great importance in the study of minerals, as it often affords a ready means of distinguishing them. The method of measuring hardness consists in selecting a series of solids, each one of the series being harder than the one above it, and softer than the one below it. At one end of the series, therefore, the hardest solid known is placed; at the other end, the softest which we may wish to measure.

DISTINCTIVE CHARACTERS OF LIQUIDS.

We have now to learn the leading properties which liquids possess, which distinguish them from solids on the one hand and from gases on the other. We have already learned that, being forms of matter, they have certain general characters in common with all other material things; but what is there about a liquid which makes us give it a name of its own? **A liquid adapts**

itself to the shape of the vessel containing it, but the conditions remaining the same, it keeps its own size or volume, however much its shape may vary. *When it is not held by the sides of a vessel it at once flows.* This is the commonest everyday experience. You cannot get a pint of beer into a glass of half a pint capacity. It does not matter what the shape of the bottle or jug may be—providing it holds a pint, as we say, or provided its capacity is a pint, the quantity of beer taken to exactly fill it is always the same. If we turn the jug upside down, the beer all runs away because there is no part of the vessel to prevent it from *flowing*. The surface of liquids, too, enclosed by a vessel is always level.

The Flow of Liquids.—The power of flowing is not perfect in liquids. The small particles making up the liquid always stick to one another a little, and when any part of a mass of liquid moves, it always attempts to drag the neighbouring particle, which is at rest, with it. We can sum this up by saying that liquids would flow perfectly if they possessed no *viscosity*. Those liquids which have little viscosity, or, what is the same thing, are very *mobile* liquids, are instanced by alcohol and water; while treacle and tar have little *mobility*, but are very *viscous*. Evidently, then, there is a gradation in those forms of matter which have as yet come before our notice. At one end we have very mobile liquids, which as the viscosity increases flow less and less easily, until at last there is no power of flowing at all, and we have the solid form of matter.

EXPT. 14.—Procure specimens of treacle and pitch. Soften the latter. Compare the consistency of the treacle and the softened pitch with that of water, and note the gradual increase in the viscosity of the liquids.

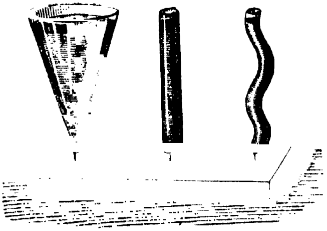

FIG. 5.—Vessels of various shapes in communication. Water standing at same level in all.

Liquids find their Level.—If several vessels of the most varied shapes (Fig. 5) are in communication with one another, and water be poured into any one of them, we shall find that as soon as the water has come to rest it will stand at the same level in all the tubes, however different the form of the vessels may be. It is this property of liquids which is utilised in the construction of the *water-level.* Its construction and use will be easily understood by a glance at Fig. 6. However the doubly-bent tube may be standing, the line joining the two surfaces will always be level. This is of great service to surveyors and such people who must be able to find a level line for the purpose of their observations. The following simple experiment is useful in showing that the surface of a liquid at rest is level:—

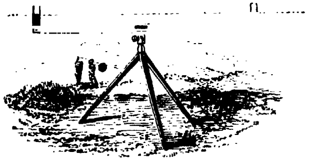

FIG. 6.—A Water-level.

EXPT. 15.—Into a shallow glass vessel pour enough mercury to cover the bottom. Attach a ball of lead to the end of a fine string, and so construct a *plumb-line.* Hang it over the surface of the mercury, and notice that the line itself and its reflection are in one and the same line. If this were not the case, that is, if the image slanted away from the plumb-line itself, we should know the surface of the liquid was not horizontal.

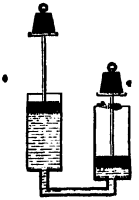

FIG. 7.—To illustrate Hydrostatic Pressure. The two cylinders are of the same size, so equal weights upon the pistons balance one another.

Liquids communicate Pressure equally in all Directions.—It will be desirable first to satisfy ourselves that liquids communicate pressure, and then try to understand the second part of the statement, that they communicate it equally in all directions. Imagine we have two cylinders in connection, as shown in Fig. 7. Into each cylinder fit a piston with a plate attached to the top, as our illustration also makes clear. If we push one piston down, we notice tha the other moves up. If we put a weight, say 10 lbs., on to each piston, they will exactly balance one another, and there is no movement in either case. Each piston is pressed upwards and downwards to an equal degree, and in consequence does not move at all. The experiment can be varied by making one cylinder horizontal, while the other remains vertical. In this case, also, if we push in the horizontal piston we notice that the upright piston rises in just the same manner as before, showing us that pressure is communicated *in different directions*. Finally, if we use a device after the pattern of our illustration Fig. 8, and fill it with water, when the piston working in the cylinder is pushed down the jets spurt out equally in all directions from the holes in the globe, thus showing that the pressure is communicated *equally in all directions*.

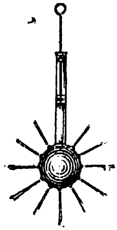

FIG. 8.—Equality of Liquid Pressure.

Hydraulic Press.—Reverting to our first example in the last paragraph, and glancing at Fig. 7, suppose that the surface of the piston on the right had been twice as great as the other, and that, as before, weights of 10 lbs. had been placed upon each piston. They will no longer balance; the right-hand weight is pushed upwards, and to bring about a balance it would be found necessary to put 20 lbs. on the larger piston. Similarly, had the right-hand piston been a hundred times larger, we should have to put 1,000 lbs. upon it to bring about a balance. The upward force, then, is proportional to the extent of the surface of the piston. This principle, which seems so different from what we should naturally expect, is referred to as the *Hydrostatic Paradox*, and is utilised in the Hydraulic Press, called, after its inventor, the *Bramah Press*, which is shown in Fig. 9. Here we have exactly the condition of things just described, two cylinders in connection, with pistons fitted into them, one much larger than the other.

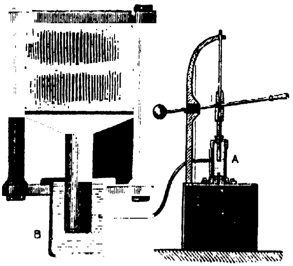

FIG. 9.—Hydraulic Press.

The application of a comparatively small force to the small one is felt on the larger one, and it is as many times greater in an upward direction as the piston of B is larger in area than the piston in A. This great upward force is being used in the instance shown in the diagram to compress bales of wool.

Liquids can be separated into Drops which will run together again.—Liquids are able to form drops because of a property, known as *cohesion*, which they possess. The larger the drop which can be formed the greater is the cohesion between the particles of the liquid.

Cohesion is the mutual attraction which the molecules of a body exert upon one another; it is, indeed, the force which keeps the particles of a substance together. Cohesion is strongest in solids, which in its absence would crumble into powder; it acts between the particles of liquids, but in gases it may said to be absent.

The following experiments beautifully show this property of liquids :—

Expt. 16.—Sprinkle some powdered resin on a board and then a little water. Notice the water collects in *drops;* the smaller they are, the more nearly spherical they are. Observe the same thing with mercury on a sheet of paper.

Expt. 17.—Mix methylated spirit and water until a few drops of oil just float when the mixture is quite cool. Pour fresh oil, by means of a pipette, into the middle of the mixture. Notice spherical globes of oil can thus be formed.

Expt. 18.—Observe that drops once formed can be made to run together again by coming in contact.

It will be as well before leaving this subject to learn in what way cohesion differs from another similar property possessed by some kinds of matter and which is known as *adhesion.* Adhesion denotes the attraction between unlike particles of matter. A metal plate may be made to *adhere* to glass ; and a postage stamp may be said to exhibit the existence of the force of adhesion when it is stuck upon an envelope.

Another experiment will make clear the difference between this *adhesion* exhibited between the postage stamp and the envelope and the force of *cohesion* :—

Expt. 19.—Carefully clean and polish two pieces of plate glass, and then place one on the other. The two surfaces will be found to *cohere*, and considerable force will be required to separate them.

Distinctive Characters of Gases

Characters of Gases.—It has already been explained that the leading difference between a solid and a liquid is the power of flowing which the latter possesses. Gases also possess fluidity, and to a much more marked degree than liquids. But whereas liquids are almost incompressible, gases are very easily compressed into a much smaller space according to a definite law, viz., just in that proportion in which you *increase* the pressure on a gas do you *decrease* the volume which it occupies. Nor are these the only differences. A liquid always adapts itself to the shape of the containing vessel, and presents a level surface at the top; a gas, on the other hand, will, however small its volume, immediately spread out and do its best to fill a vessel, however large; and it does not present any surface to the surrounding air. We can never say exactly where the gas leaves off and the air begins. Another distinction will be more fully appreciated after we have considered the action of heat upon the volume of bodies. We shall learn that, generally speaking, all bodies get larger as they are heated; this is very much more decidedly the case with gases than with liquids. **Gases,** then, **are easily compressible and expand indefinitely.**

We shall learn that gases expand equally when heated to the ame extent, but this fact and others will be much better undertood in their proper places.

Constancy of Weight in different States.—When a solid is onverted into a liquid, or a liquid into a vapour, no change of weight is experienced. This has been found to be true in all cases, and the following experiments will illustrate the fact:—

Expt. 20.—Boil water in a flask or a retort, as in Fig. 10, and catch the condensed steam, taking care that none escapes, in another flask kept cool by resting in water. The water thus collected will be found to have the same weight as that boiled away.

Expt. 21.—Place a piece of ice in a flask suspended from one arm of a balance. Counterpoise the flask with the ice in it; then melt the ice by warming the flask, and show that the counterpoise is unaltered.

Expt. 22.—Put some warm water in a flask and some salt in a piece of paper. Counterpoise the flask of water and the paper of salt together and then dissolve the salt in the water. The total weight remains unaltered.

No kind of Matter can be destroyed.—There is a certain fixed amount of matter in the universe which never gets any less and never any greater. If we confine our attention to the earth, we cannot say that it never receives an addition to the matter of which it is built, for every year it is receiving numbers of small solid bodies which are continually falling upon its surface from outside space. But the proposition means that in those cases in which it is popularly supposed there is a loss of matter, for instance when a fire burns out, no such destruction has taken place, but only a change in the form assumed by the matter. It will make the statement quite clear if we follow out what really takes place when a candle burns, and, as it would seem, gradually disappears.

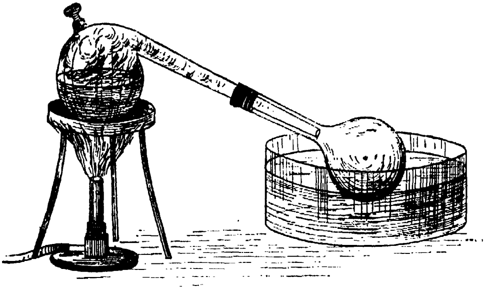

Fig. 10.—Condensation of Steam.

Expt. 23.—Over a burning candle hold a white glass bottl which has been carefully dried inside and out. Observe tha the inside of the bottle becomes covered with mist and after short time drops of liquid are formed which run down th sides of the bottle. *The burning of the candle has resulted in the formation of a new form of matter.*

EXPT. 24.—Allow the candle to burn in a similar bottle placed on the table. After a time the candle ceases to burn, and when this has happened take the candle out and cover the bottle over with a glass plate. Notice that no change seems to have taken place in the gas which filled the bottle. Now pour some clear lime-water into another clean bottle and shake it up, the lime-water remains clear.[1] Lift off the glass plate and do the same with the bottle in which the candle has been burnt, the lime-water turns milky. *The burning of the candle has also resulted in the formation of a new kind of matter, viz., a gas which turns clear lime-water cloudy.*

Evidently, then, when a candle burns it ceases to exist as tallow or wax, or whatever the candle is made of, and assumes new forms, still material, one liquid, the other a gas which turns lime-water milky. If we were to weigh all the liquid formed and all the gas which turns lime-water milky, we should find that these two things together actually weigh *more* than the part of the candle which has disappeared did. The reason why there is an *increase* of weight will be explained later. The arrangement for performing this experiment is shown in Fig. 11. The candle is burned in a wide tube, A, fitted with a cork at the bottom with holes in it to allow the air, which is necessary to help the candle to burn, to pass in as shown by a bent arrow. The tube B is filled with a substance which has the power of arresting the products of the burning. Such a substance is caustic soda, which is used in the form of lumps. Air is drawn through the apparatus as shown by the arrows.

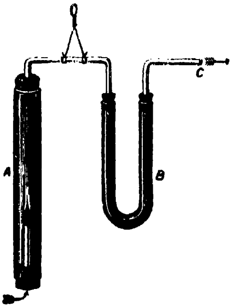

FIG. 11.—Apparatus to show that no matter is destroyed when a candle burns.

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[1] There may be a slight milkiness owing to the presence in the air of the gas which is formed. If so the experiment must be made comparative.

Before the experiment is started, the candle is weighed and also the tube containing the lumps of caustic soda. After the experiment the same things are weighed over again. It will be found that the tube B has increased in weight to a greater extent than that of the candle has been diminished. We are quite sure, therefore, that there has been no loss of matter.

Chemists have satisfied themselves that this is universally true, and it must be remembered as a truth of the highest importance, **matter is indestructible.**

CHIEF POINTS OF CHAPTER I.

Matter.—When we speak of matter we mean all things which exist in or out of our world, which we can become aware of by the help of our senses.

Properties are certain effects caused by the things which are said to possess them.

Properties possessed by Matter.—Matter occupies space, offers resistance, possesses weight, and transfers motion to other things when it strikes against them.

The same Matter can exist in three different States.—The three states of matter are the **solid, liquid** and **gaseous.** By suitable means, as in the experiments described in the chapter, the same portion of matter can be made to assume these states in order. Sometimes the change from one state to another is gradual, sometimes sudden. There is no hard and fast line between the three conditions of matter.

Distinctive Characters of Solids.—A solid body does not readily alter its size or shape. It will keep its own volume and the same form unless subjected to a considerable force. Or, we may say solids possess rigidity.

Solids possess Elasticity, Tenacity, Ductility and Hardness.

Elasticity is the tendency to go back to the original form or volume after being forced out of it.

Tenacity is measured by ascertaining what weight is necessary to break solids when in the form of wires.

Ductility is the property by virtue of which solids can be made into wires.

Malleability is a similar property to ductility which enables certain solids to be beaten out into sheets.

Hardness is the property by virtue of which solids offer resistance to being scratched or worn by others.

Distinctive Characters of Liquids.—A liquid adapts itself to the shape of the vessel containing it, but the conditions remaining the same, it keeps its own size or volume, however much its shape may vary.

Liquids possess Fluidity.—When a liquid is not held by the sides of a vessel, it at once flows. This power of flowing is not perfect in liquids, or we may say, liquids are not perfectly mobile. They all possess a certain degree of *viscosity*.

Other Properties of Liquids.—They find their level. They communicate pressure equally in all directions. They can be separated into drops which will run together again.

Distinctive Characters of Gases.—They possess fluidity to a much more marked degree than liquids. Unlike liquids they can easily be compressed into a much smaller space. A gas, however small its volume, will spread out and do its best to fill a vessel, however large. Expressed shortly we say, gases are easily compressible and expand indefinitely.

Constancy of Weight in different States.—When a solid is converted into a liquid, or a liquid into a vapour, no change of weight is experienced.

Matter is Indestructible.—There is a certain fixed amount of matter in the universe which never gets any less and never any greater. Whatever changes may occur in the composition of matter there is never any loss of weight.

QUESTIONS ON CHAPTER I.

1. What do you understand by "matter," or, a "material thing"?

2. Give some of the properties which are possessed by all kinds of matter and explain in your own words what is meant by a property.

3. Describe experiments which prove :—
   (*a*) That solids are porous.
   (*b*) That liquids, too, contain pores.

4. What experiment could you perform to show that a solid, say a billiard ball, is elastic? Explain as well as you can what you mean by elasticity.

5. The same portion of matter can, under suitable conditions, assume different states. Describe fully some experiment which illustrates this statement.

6. What evidence can you give that the different states of matter gradually shade into one another?

7. What properties are generally associated with matter in the solid form? Give a definition of a solid which includes the chief of these.

8. In what respects are liquids different from solids?

9. How do gases and liquids differ?

10. What property in particular is possessed by liquids and not by solids? And what character has a gas which neither liquids nor solids possess?

11. What property is it which liquids possess which enables them to form drops? Describe another experiment which also shows the possession of this property by liquids.

12. Give in the form of a definition the distinctive properties of a

liquid. What name do we give to the property which prevents liquids from being perfectly fluid and what do you know about it?

13. Describe experiments in support of the assertion that liquids communicate pressure equally in all directions.

14. What is meant by the "hydrostatic paradox"? Describe some useful apparatus in which this principle is utilised.

15. How would you show by experiment that the weight of the same portion of matter in different states is constant.

16. On what grounds is the assertion that "Matter is indestructible" based?

CHAPTER II

UNITS OF LENGTH, AREA AND VOLUME

Measurement of Length.—We have from time to time in the preceding chapter referred to lengths as being of so many yards, or feet, or centimetres, and it is desirable, before going further, to acquaint the student with the exact significance of these and other measures of length. It is clear that before comparing any one length with any other we must have some standard to which we can refer them. In this country the standard adopted is the length between two marks on a platinum bar kept at the Exchequer Chambers, the bar being at a certain fixed temperature when the measurement is made. This length is quite arbitrary and is called a **yard**. The yard is subdivided into three equal parts, each of which is a **foot**. The foot is in its turn divided into twelve equal parts, called **inches**. Multiples of the yard are also used and special names given to them, thus:—

2 yards = 1 fathom.
5½ yards = 1 rod, pole, or perch.
40 poles = 1 furlong.
$\left.\begin{array}{r}8 \text{ furlongs}\\ 1{,}760 \text{ yards}\end{array}\right\} = 1 \text{ mile.}$

We have given these in full to show how clumsy and unsatisfactory the British measures of length are, and to point out that it is for this reason they are not used in scientific work, even by British men of science.

French geometricians decided that such an arbitrary standard was, in view of the chance of its loss or destruction, an undesirable one, and suggested that if a fraction of the circumference of the earth were taken they would, in the event of the loss of the standard, be easily able to replace it by an exact copy. They proposed the one ten-millionth part of the earth's quadrant, *i.e.*, of the distance from the equator to the pole, as a suitable length, and this they called the **metre**. After bars had been prepared of this length it was unfortunately found that the length of the quadrant had not been exactly determined, and consequently the length of the standard metre at Paris is arbitrary, and we must define the standard metre as being the length at a certain temperature between two marks on a platinum bar kept at Paris. It is equal to 39·37079 inches. The metre is subdivided into ten equal parts, each of which is called a **decimetre,** the tenth part of the decimetre is called a **centimetre,** and the tenth part of the centimetre is known as a **millimetre.** Thus we get

10 millimetres = 1 centimetre.
10 centimetres } = 1 decimetre.
100 millimetres }
10 decimetres }
100 centimetres } = 1 metre.
1,000 millimetres }

The multiples of the metre are named **deka-, hecto-,** and **kilometres.** Their value is seen from the following table :—

10 metres = 1 dekametre
100 metres = 1 hectometre
1,000 metres = 1 kilometre

The kilometre is equal to about three-fifths of a mile.

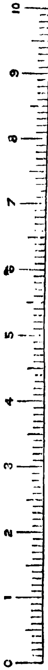


FIG. 12.—A Decimetre subdivided into centimetres and millimetres.

Exercises in Linear Measurement.—EXPT. 25.—Procure a rule divided into inches and parts of an inch on one edge, and centimetres and parts of a centimetre on the other. Measure the length of this page both in inches and centimetres; also determine other lengths in the two systems of measurement.

Put down the results in parallel columns, as shown below, and from them calculate the number of centimetres in an inch.

Length in centimetres.	Length in inches.	$\frac{\text{No. of centimetres.}}{\text{No. of inches.}}$

Expt. 26.—Draw a line a metre long and one a yard long close to it. Measure the length of the metre line in inches and fractions of an inch, and the length of the yard line in centimetres. In this way find

Number of inches in one metre = 39·3.
Number of centimetres in one yard = 91·4.

Expt. 27.—Draw a line one foot long and find its length in centimetres; draw another a decimetre long and find its length in inches. In this way determine

Number of centimetres in one foot = 30·5.
Number of inches in one decimetre = 3·9.

Expt. 28.—Measure the length of your desk, or of a table, or other convenient object, in inches and decimals (tenths) of an inch. Determine the same length in metres and decimals of a metre. Use the results to determine the number of inches in a metre, thus:—

Length in inches.	Length in metres.	$\frac{\text{No. of inches.}}{\text{No. of metres.}}$

Relation between British and Metric Units of Length.—The exact relations between linear measures in the British and Metric systems of units are shown in the following tables, and also in Fig. 13.

Metric to British.

1 millimetre = 0·039 inch.
1 centimetre = 0·394 inch.
1 decimetre = 3·937 inches.
1 metre = 39·371 inches.
1 metre = 3·281 feet.
1 metre = 1·094 yards.
1 kilometre = 0·621 mile.

British to Metric.

1 inch	=	25·399 millimetres.
1 foot	=	0·305 metre.
1 yard	=	0·914 metre.
1 mile	=	1609·000 metres.
	=	1·609 kilometres.

Simple application of Measurements of Length.—EXPT. 29.—Measure the circumference of a glass bottle by winding a strip of paper around the curved surface and marking the point where the edge meets the commencing corner. Determine the length of the paper from the corner to the mark by measuring it with a centimetre scale. Next measure carefully the diameter of the bottle. Repeat the experiment with other objects, such as a cylindrical canister or a wooden cylinder or disc. Set down the results in parallel columns and use them to find the ratio of the circumference of a circle to the diameter.

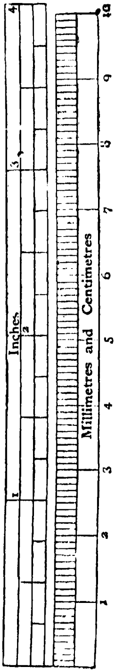


FIG. 13.—Comparison of Inches and Centimetres.

Diameter.	Circumference.	$\frac{\text{Circumference.}}{\text{Diameter.}}$

Measurement of Area.—In describing the measures of area (or space of two dimensions) according to the English system we use the same names as when measures of length are referred to, simply prefixing the word *square*. Thus we speak of square inches and square feet, and since there are 12 linear inches in a linear foot, there will be 12 × 12 = 144 square inches in a square foot, and similarly throughout the measure. In the same way it is the custom to speak of an area, as of so many square centimetres, or square metres, as the case may be.

British Units of Area.

144 square inches = 1 square foot.
9 square feet = 1 square yard.
$30\frac{1}{4}$ square yards = 1 square rod, pole, or perch.
40 square poles = 1 rood.
4 roods = 1 acre.
640 acres = 1 square mile.

Units of Area in the Metric System.

100 square millimetres = 1 square centimetre.
10,000 square millimetres, 100 square centimetres = 1 square decimetre.
1,000,000 square millimetres, 10,000 square centimetres, 100 square decimetres = 1 square metre. = 1 centiare.
10 centiares = 1 deciare.
10 deciares = 1 are = 1 square dekametre.
10 ares = 1 dekare.
10 dekares = 1 hectare = 1 square hectometre.

Exercises in Measurement of Areas.—EXPT. 30.—Draw several right-angled triangles (as in Fig. 14) and carefully measure the lengths of the sides of each of them.

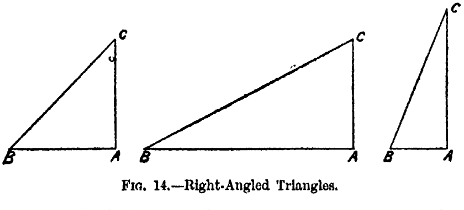


FIG. 14.—Right-Angled Triangles.

Square each of these numbers and set down the results in parallel columns, as here shown :—

$(AB)^2$.	$(AC)^2$.	$(AB)^2 + (AC)^2$.	$(BC)^2$.

This is not only an exercise in measurement, but the results will show that the square of the longest side of a right-angled triangle is equal to the squares of the two other sides added together.

EXPT. 31.—Draw a square inch and a square centimetre, side by side, as shown in Fig. 15.

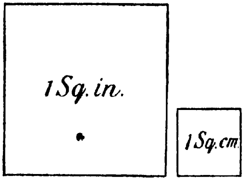


FIG. 15.—Square Inch and Square Centimetre.

EXPT. 32.—Draw a square decimetre and divide it into square centimetres, as shown on a reduced scale in Fig. 16. The area of the square is thus shown graphically to be equal to the length multiplied by the height. Now draw two or three oblongs and determine their areas by means of this rule. Determine the areas of the square and oblongs both in square inches and square centimetres, and use your results to find the number of square centimetres in one square inch, thus :—

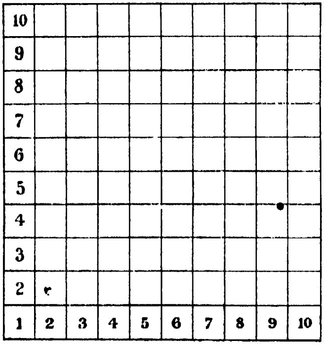


FIG. 16.—Square divided into Units of Area.

Area of a given rectangle in square inches.	Area of same rectangle in square centimetres.	Square centimetres. Square inches.

EXPT. 33.—Draw a square foot and a square decimetre and find the number of square decimetres in the square foot, and the fraction which one square decimetre is of one square foot.

EXPT. 34.—If convenient, draw a square yard and a square metre upon a blackboard or a large sheet of paper. Determine the area of the square yard in decimals of a square metre, and the area of the square metre in square yards. To obtain the latter result, the number of square inches in one square metre is determined, and the result divided by the number of square inches in a square yard.

Relation between British and Metric Units of Area.—These exercises have shown roughly the relation between the measurements of areas expressed in the metric and British systems of units. The exact proportions which the units of area in the two systems bear to one another are shown in the following table:—

Metric to British.

1 square centimetre = 0·155 square inch.
1 square decimetre = 15·500 square inches.
1 square metre = 10·764 square feet.
= 1·196 square yards.
1 are (100 square metres) = 119·603 square yards.
1 hectare (10,000 square metres) = 2·471 acres.

British to Metric.

1 square inch = 6·451 square centimetres.
1 square foot = 9·289 square decimetres.
1 square yard = 0·836 square metres.
1 acre = 10·117 ares.
1 square mile = 258·989 hectares.

Determinations of Areas.—It has already been explained that the area of a square or an oblong can be determined by dividing the figure into units of area and counting the number of units. Paper can be procured divided into squares of a definite size, similar to those represented in Fig. 17. By

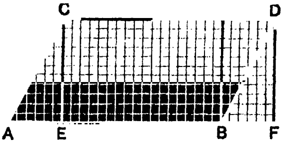


FIG. 17.—The Area of the Parallelogram CABD is the same as that of the Oblong CEFD.

means of "squared" paper of this kind, the area of any figure can be found by drawing the figure upon the paper and counting the number of squares embraced by the boundary lines. If "squared" paper is not available, the area is still easily found by dividing it up into square centimetres with a T-square and set squares.

EXPT. 35.—Draw an oblong, such as CEFD (Fig. 17), upon squared paper and find the number of squares it contains. Draw a parallelogram CABD with a base equal to the base of the oblong, and a height equal to that of the oblong. Count the number of squares in the parallelogram (counting the portions of squares as quarters, halves, and three-quarters). The number will be the same as that found for the oblong. No matter what parallelogram is drawn, the area will be found to be the same as that of an oblong on the same base and having the same height.

EXPT. 36.—Draw a parallelogram on paper or thin card and then cut it in two from corner to corner. You have now two triangles, and by laying one on the other you will find that they fit and are equal.

EXPT. 37.—Draw any irregular figure upon squared paper and count the number of squares embraced by it. If the size of the squares is known, the area of the figure can be determined in this way. The area of any irregular figure can evidently be determined by tracing the figure upon squared paper and counting the number of squares included by the outline, or by actually dividing it into square centimetres as shown above.

Measurement of Volume.—When we come to measure volumes we are dealing with three dimensional spaces; and just as a plane surface or area measuring one foot in each of the directions, length and breadth, is called a square foot from the name of the figure which it forms, so a solid which is obtained by measuring a foot in three directions, at right angles, length, breadth, and thickness, is called a cubic foot, from the name *cube* given to the solid so formed. Similarly using the metric system, we may speak of a cubic metre, or a cubic decimetre.

The amount of space enclosed by a solid figure is called its *volume.* The volume of a solid is the space it occupies or its size. The volume of a vessel is the amount of space it encloses.

In the metric system a special name is given to the volume of a cubic decimetre, that is, a cube having a decimetre edge. It is called a **litre**, and is equal to about one and three-quarters English pints. The sub-multiples and multiples of a litre are named in a similar way to those of the metre. There is no such simple relation between the measures of length and volume in the English system, though the gallon is defined as a measure which shall contain 10 lbs. of pure water at a certain temperature and pressure.

British Units of Cubic Measure or Volume.

1728 (= 12 × 12 × 12) cubic inches = 1 cubic foot.
27 (= 3 × 3 × 3) cubic feet = 1 cubic yard.

Units of Volume in the Metric System.

10 Centi-litres = 1 deci-litre.
10 deci-litres = 1 **litre** = 1 cubic decimetre.
10 litres = 1 deka-litre.
10 deka-litres = 1 hecto-litre.
10 hecto-litres = 1 kilo-litre = 1 cubic metre.

Exercises in the Measurement of Volumes.—Expt. 38.—Cut out a cubic centimetre and a cubic inch from a piece of soap, or procure them made in wood. If possible, procure also a cubic decimetre divided into centimetre cubes like the cubic foot shown in Fig. 18. Place one cube alone on a table; then add nine more to it, so as to make a row of 10 cubic centimetres (10 c. cm.). Place 10 such rows upon one another and there are 100 c. cm.; and 10 such slabs of 100 c. cm. each make altogether 1,000 c. cm. One cubic decimetre is thus shown to contain 1,000 c. cm.

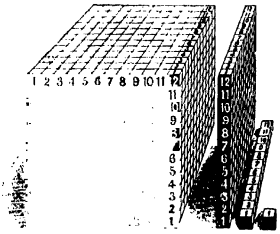


Fig. 18.—Showing a small cube, a row of 12 cubes, a slab of 12 rows, and a cube of 12 slabs.

The exercise also demonstrates that volume involves length, height, and breadth.

Expt. 39.—Construct a cubical box of a decimetre edge out of cardboard, and make it water-tight by varnishing. The capacity or volume of such a box is a cubic decimetre or a litre. Make similar boxes 1 cubic centimetre and 1 cubic inch in capacity.

Expt. 40.—Procure a half-pint glass measure graduated into fluid ounces, and one graduated into cubic centimetres. Use these glasses to determine the number of c. cm. in a half-pint, and the number of fluid ounces which are equal to 1,000 c. cm., that is, 1 litre.

Expt. 41.—Fill the cubic inch box, from Expt. 39, with water, and by pouring the water into the jar graduated into cubic centimetres, determine the number of c. cm. in a cubic inch.

Relation between British and Metric Units of Volume.—These exercises approximately show the relations between the units of volume or capacity in the British and metric systems of measurement. The exact equivalents are as follows :—

Metric to British.

1 cubic centimetre = 0·061 cubic inch.
1 cubic decimetre = 61·027 cubic inches.
1 cubic metre = 35·316 cubic feet.
= 1·308 cubic yards.
1 litre = 1·761 pints.

British to Metric.

1 cubic inch = 16·386 cubic centimetres.
1 cubic foot = 0·028 cubic metre.
= 28·315 cubic decimetres.
1 cubic yard = 0·764 cubic metre.
1 pint = 0·568 litre.
1 quart = 1·136 litres.
1 gallon = 4·543 litres.

Simple Application of Volume Measurement.—The student cannot have failed to notice that the

volume of a cube	=	area of the base	×	height
(in, *e.g.*, c. cm. or c. inches).		(in, *e.g.*, sq. cm. or sq. inches).		(in, *e.g.*, cm. or inches).

The volume of any rectangular block, or of a cylinder, can be calculated by the application of the same rule. This can be easily understood by referring to the accompanying diagrams (Fig. 19). The base of each block illustrated is divided into units of area—say, square centimetres. If we consider the cube to be divided into slices one centimetre thick, each slice could be divided into cubic centimetres, and the number of cubic centimetres would be the same as the number of square centimetres in the base. Hence, if the base had an area of 30 square centimetres, the slab 1 centimetre high would contain 30 cubic centimetres. Two such slabs would therefore contain 60 cubic centimetres, three slabs 90 cubic centimetres, and so on for any number of slabs. The volume of any block having the same width and breadth all the way up may evidently be reasoned out in the same manner, being equal to the area of the base multiplied by the vertical height.

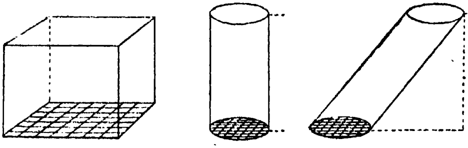


FIG. 19.—Volumes of Regular Solids.

To determine the volume of an irregular solid, we proceed by an indirect method as follows :—

EXPT. 42.—Fill the cubic inch box with water and pour the water into a narrow glass jar or bottle. Make a mark upon the bottle level with the top of the water. Repeat the operation until the vessel is full of water. A bottle graduated into cubic inches is thus obtained, and it can be used to determine roughly the volume of an irregular solid such as a piece of lead, a glass stopper, a few nails, or any solid heavier than water which will go into the bottle. Fill the bottle about three-quarters full of water, observe the level of the water, and then gently drop in the chosen object. From the rise of water-level which immediately takes place, the volume of the object can be estimated.

EXPT. 43.—Repeat the last experiment, substituting the jar graduated into c. cm. The volume of the solid used, in metric units, is shown by the rise of the water above its original level.

If the object is lighter than the liquid employed, a sinker made of lead or iron should be employed. First find the volume of the sinker as in Expt. 42; then take the sinker out of the water and tie it with thread to the light object, and place them both in the jar. The additional rise of level which takes place will give the volume of the object itself.

Chief Points of Chapter II.

Measurement of Length.—In the British system the standard of length is the *yard*. It may be defined as the length between two marks on a platinum bar kept at the Exchequer Chambers, the bar being at a certain fixed temperature when the measurement is made. The yard is subdivided into *feet* and *inches*. The multiples of the yard are given on p. 21.

In the metric system the standard is the *metre*; it may be defined as the length, at a certain temperature, between two marks on a platinum bar kept at Paris. It is equal to 39·37079 inches. It is subdivided into *decimetres*, *centimetres* and *millimetres*. The multiples are the *dekametre*, *hectometre* and *kilometre* (p. 22).

Measurement of Area.—In measuring areas or space of two dimensions, the same names are employed as when speaking of lengths with the word *square* prefixed. We thus get such terms as *square foot* and *square decimetre*. Since there are 12 linear inches in a linear foot there will be 12 × 12 = 144 square inches in a square foot. The British and metric units of area are given on p. 25.

Measurement of Volume.—In measuring volumes, or space of three dimensions, we employ, in the British system, such terms as *cubic foot* and *cubic yard*. Since there are 12 linear inches in a linear foot there will be 12 × 12 × 12 = 1728 cubic inches in a cubic foot.

In the metric system a special name is given to a *cubic decimetre*, that is to the volume contained by a cube having a decimetre edge, it is called **a litre** and is equal to about one and three-quarters English pints. The names of the multiples and sub-multiples of the litre are formed in the same way as in the case of the metre (p. 29).

Relation between British and Metric Units in the cases of length, area and volume are given in the course of the chapter and should be carefully examined by the student.

Questions on Chapter II.

1. What is the necessity for units of length, area and volume? What is the unit of length in common use in this country and how may it be defined?

2. What is the unit of length in the metric system? What rela-

tion was it originally meant to bear to the earth's circumference? How would you define the metric unit of length in common use?

3. What advantages do you associate with the use of the metric system?

4. Give the multiples and sub-multiples of the unit of length in the British system. How are they related to one another?

5. What names are given to the multiples and sub-multiples of the metre? Compare the metre and the yard in terms of one another.

6. Compare and contrast the British and metric systems of area and volume.

7. Explain how the metric units of volume and length are related. Is there any such simple relation in the case of British units?

8. Describe an experiment for determining the volume of a solid heavier than water.

9. How would you proceed to find out the volume of a cork or a deal cube?

10. Write down the length of a metre in yards, feet and inches. How many centimetres are there in a metre, and how many millimetres in a centimetre.

11. How many millimetres and centimetres are equal to the length of an inch.

CHAPTER III

UNITS OF TIME, VELOCITY AND ACCELERATION

The Earth's Rotation.—The apparent daily motion of the sun and stars across the sky is a direct consequence of the earth's rotation on its axis. The sun appears to regularly go through certain periodic changes of position. It rises, travels higher and higher into the sky, reaches its highest position, sinks lower and lower, and finally sets. When the sun is at its highest altitude on any day it is due south, and is said to *south* or be *southing*. The interval of time between the sun's highest position on any one day to its corresponding position on the next succeeding day is an **apparent solar day.** These apparent solar days vary in length throughout the year.

Expt. 44.—Fasten a small rod at right angles to a flat board. Place the board flat on a table so that the rod is vertical. Move a candle in a semicircle above the table; and note the change in the angle that the shadow of the rod makes. Compare the conditions of the experiment with the measurement of the solar day by means of a sun-dial.

Mean Solar Day.—It has been pointed out that the length of days measured by the sun varies throughout the year, hence no single one of these days will do for a convenient standard of time. But if the lengths of all the days in the year be added together, or the length of a year measured by the sun be divided by the number of days in the year, we obtain an interval of time which is always the same. Such a day, which is of course an imaginary one, is called a **mean solar day.** Sometimes the mean solar day will be longer than the solar day, sometimes it will be shorter, and occasionally both days will be of exactly the same length.

Solar time is known as *apparent time*, and clock time as *mean time*.

Sidereal Day.—Just as in the case of the sun so with all the stars, they rise, south, and set. But whereas with the sun the interval between two successive southings varies throughout the year, it is found that the time which elapses between two succeeding southings of a star at any season of the year is always the same. This interval constitutes **a star** or **sidereal day.**

Period of Rotation of the Earth.—As the apparent motions of stars across the sky are produced by the rotation of the earth, it will be evident that the exact time of rotation can be determined by finding the interval which elapses between two successive returns of any particular star to the same point of the sky. A star may, indeed, be regarded as a fixed reference mark under which the earth turns; so that by observation of it we are able to determine the time taken by the earth to spin round once. The interval between two successive transits of the same star, or, as it is called, a sidereal day, is the time of such rotation.

No matter what star is selected for observation the interval is the same, thus showing that the earth is a rigid body, and that all parts of its surface have the same angular velocity.

Units of Time.—The sidereal day, like the mean solar day, is subdivided into hours, minutes, and seconds, but as the latter is four minutes longer than the former, the units are not of the same value. We may take either the *mean solar second* as the unit of time, or the *sidereal second*. In the former case the unit is founded on the average length of the solar day, and in the latter upon the length of the invariable star day, or the time of rotation of the earth upon its axis. But in either case the second, that is, the unit of time, is the 86,400th part of the day used.

In physical measurements the unit of time adopted is the mean solar second, that is, it is founded on the average time required by the earth to make one complete rotation on its axis relatively to the sun considered as a fixed point of reference.

Instruments for Measuring Time.—We need only concern ourselves with the modern contrivances for measuring time, viz., clocks and watches. It will be sufficient to regard these as instruments for measuring intervals of time in terms of the mean solar day to which attention has been directed. In a clock the rate is regulated by means of the pendulum, the properties of which can be best understood by an experiment.

> EXPT. 45.—Attach a weight to the end of a cord. Fix the cord in such a way that the pendulum can oscillate freely. Set it oscillating, and notice how long it takes for the pendulum to complete a given number, say twelve, swings. Keeping the cord exactly the same length, attach a heavier weight and repeat the experiment. The time of swing remains unaltered. Keeping any one weight, observe the time taken to complete twelve swings when the length of the cord is varied. It will be seen that the time of swing varies with the length of the cord. Notice also that it does not matter if the pendulum makes a wide oscillation or a very small one, the time taken being the same.

If it were possible for the student to perform the experiment, it would be found that the time taken for the pendulum to swing backwards and forwards varies as it is taken from the equator to the poles, on account of the fact that the earth is not exactly spherical in shape. Or, putting the same fact in another way, in order that a pendulum may swing backwards and forwards in the same interval of time, it is necessary to alter the length of the cord in our experiment as we travel from the equator towards either pole. A pendulum of such a length that the distance from the point of suspension to the centre of the bob is 39·139 inches, swinging at Greenwich, completes one swing in a second of time. In a clock we have a mechanical contrivance for maintaining the swinging of a pendulum. We must content ourselves with referring the reader to books on astronomy and horology for an account of the construction of a clock. In watches the place of the pendulum is taken by a carefully suspended balance wheel.

Definition of Motion.—The word *motion* is meant to convey the idea of *change of place*. The simplest forms of motion are changes in the positions of bodies with regard to one another. When a boy runs down the street he is in *motion ;* as regards the houses and lamp-posts he moves. To fully describe the boy's motion it would be necessary to know the *direction* in which he is moving or the *line* along which he runs, and the *rate* or *velocity* with which he travels.

Velocity.—Velocity is, then, **the rate at which a body moves,**

or the number of units of length it moves over in a unit of time. Velocity may be either *uniform or variable.* In the case of our boy, if during every second through which he moves he travels over a line of five yards in length, we should say he had a *uniform* velocity of five yards a second. But suppose this boy does not move regularly over five yards in every second; he sometimes dawdles, sometimes stops to look at a shop, at other times he puts on a spurt to make up for lost time. How should we describe his motion now? His rate varies from time to time, or his velocity is *variable*, and to describe such a variable velocity it is usual to speak of the velocity *at any instant* as being a certain number of yards per second. Say the boy of our example, moving with a variable velocity, has at a given instant a velocity of eight yards per second. We should mean that he would, if he continued to move at the same rate as he had at the given instant, travel over eight yards in the succeeding second.

Average Velocity.—But it is sometimes better to find the *average velocity* of the moving body. Returning to our boy, suppose he travelled 800 yards in 400 seconds; if we divide the first number by the second we obtain the boy's average rate, namely, two yards in a second; this, then, is the rate with which he would have had to travel, if he moved uniformly, in order to complete his journey in the same time.

The **unit of velocity** is the velocity of a point which passes over the unit of length in a unit of time; it is generally taken as being a velocity of one foot per second. Thus a velocity of six means a velocity of six feet per second.

Measurement of Uniform Linear Velocity.—It is a very simple matter to calculate the velocity of a body moving uniformly in a straight line when we know the distance it has travelled, measured in units of length, and the time it has taken to perform the journey, measured in units of time. Thus if we represent the space it has passed over, measured in units of length, by the letter s, and the time taken, measured in units of time, by the letter t, all that we have to do in order to find its uniform velocity is to divide the number of units of length passed over by the number of units of time taken to complete the distance, or—

$$\text{uniform velocity } (v) = \frac{\text{no. of units of length passed over}}{\text{no. of units of time taken}} = \frac{s}{t}$$

or

$$\mathbf{v} = \frac{\mathbf{s}}{\mathbf{t}}.$$

EXAMPLE.—A train, travelling at a uniform rate, completes a distance of 10 miles in 15 minutes; what is its velocity in feet per second?

Answer: 10 miles $= 10 \times 1{,}760 \times 3$ feet $= 52{,}800$ feet
15 minutes $= 15 \times 60 = 900$ seconds

$$\therefore \text{ since } \quad v = \frac{s}{t}$$

$$v = \frac{52{,}800}{900} = 58\frac{2}{3} \text{ feet per second,}$$

or the train travels at a uniform velocity of $58\frac{2}{3}$ feet in every second.

Acceleration.—An express train starting from a terminus begins to move slowly, and, as the journey proceeds, the rate of motion goes on increasing until it gets its full speed. A stone let fall from a height similarly starts from rest, and as it moves it goes faster and faster until brought to a standstill again on reaching the ground. Or we might imagine a cyclist starting for a run, and regularly increasing his speed until he could not go any faster. In all these examples the velocity of the moving body has regularly increased and the rate at which the change has taken place is spoken of as *acceleration.*

Acceleration is the Rate of Change of Velocity.—But it may be of an exactly opposite kind to the instances given above. Reverse each of the examples and consider what happens. An express train going at full speed approaches a station and its velocity is regularly diminished until it is brought to rest at the platform. A stone is thrown upwards with a certain velocity, it moves more slowly and more slowly until it comes to rest, and then starts falling. A cyclist travelling at full speed slackens his rate regularly until he comes to a standstill. In all these cases we have examples of an acceleration of an exactly opposite kind to the previous instances, but yet an acceleration. In ordinary language this kind of acceleration is given a name of its own, *retardation.*

Measurement of Uniform Acceleration.—In measuring a regular or *uniform acceleration*, we must know what addition to or subtraction from the velocity of the moving body there has been during each second of its journey. Suppose there is an addition of one foot per second to the velocity of a moving body, and that it has taken one second to bring about this change, we should refer to this as an acceleration of one foot per second in a second, or *one foot per second per second*. An acceleration which increases the velocity is referred to as *positive*, while that which diminishes it is *negative*. The first examples given above are instances of positive acceleration, while when we reverse them they afford cases of negative acceleration.

Unit of Acceleration.—As in every other measurement so, when we wish to measure accelerations, we must have a unit in terms of which we can express the quantity under consideration. **The unit of acceleration is the increase of unit velocity in a unit of time**; it is generally taken as equal to *an increase of velocity of one foot per second per second*. An acceleration of *two units* would thus be an increase of velocity of *two* feet per second per second, and, similarly, an acceleration of three units equals an increase of velocity of three feet per second per second, an acceleration of a units equals an increase of velocity of a feet per second, in one second.

If we wish to determine the velocity at any instant of a body which is moving with a uniformly accelerated velocity we must argue as follows:—Let the acceleration be a, which means an increase of velocity of a feet per second in one second. Suppose a body starts from rest, at the end of the first second it has a velocity of a feet per second, at the end of the next second $2a$, at the end of t seconds at feet per second. Or if v = change of velocity in t seconds, we can write

$$\mathbf{v = at.}$$

Consequently, to find the velocity at any instant of a uniformly accelerated moving body, all we have to do is to multiply the number of units of acceleration by the number of units of time for which it has been moving.

EXAMPLE.—An engine moves from rest with a uniform acceleration of 2 feet per second in every second. What is its velocity at the end of 5 minutes?

$$5 \text{ minutes} = 5 \times 60 = 300 \text{ secs.}$$

$$\text{Here } a = 2\,;\ t = 300\,;\ v = ?$$

$$\text{But } v = at$$

$$= 2 \times 300 = 600 \text{ feet per second.}$$

If the uniform acceleration is applied to a body already moving with a uniform velocity, it is not difficult to see that the formula or expression we have just given will become modified. Imagine a body moving with a uniform velocity of u units per second to have this velocity uniformly accelerated. Let the acceleration be a units, as in the case of the body moving from rest. Then, at the end of the first second during which the acceleration acts its velocity will be $u + a$, at the end of the second second it will be $u + 2a$, at the end of the third $u + 3a$, and at the end of t seconds $u + ta$ or $u + at$, or the formula becomes

$$\mathbf{v = u + at.}$$

EXAMPLE.—What will be the velocity of the engine in the last example at the end of the sixth minute, if it continues to move with the same acceleration?

At the end of fifth minute its velocity $u = 600$ feet per second.

$$\begin{aligned} \text{And} \quad v &= u + at \\ &= 600 + (2 \times 60) \\ &= 720 \text{ feet per second.} \end{aligned}$$

Space traversed by a Body moving under the influence of a Constant Acceleration.—The space travelled over by a body in one second is equal to its average velocity, and that travelled over in t seconds is equal to its average velocity multiplied by t. If it starts from rest and travels for one second finishing with a velocity of v feet per second, its average velocity is $\frac{1}{2}v$ during this time, and

$$\therefore \qquad s = \tfrac{1}{2}vt$$

Substituting value of v from the equation $v = at$, we get

$$\mathbf{s} = \tfrac{1}{2}at \times t = \tfrac{1}{2}\mathbf{at^2},$$

or since $t = \frac{\ }{a}$ we can write the equation $s = \frac{1}{2}vt$ thus—

$$s = \tfrac{1}{2}v \times \frac{v}{a} = \tfrac{1}{2}\,\frac{v^2}{a}, \text{ from which}$$

$$\mathbf{v^2 = 2as.}$$

Expressed in words, the rule for finding the space travelled in a given number of seconds is—multiply half the acceleration (expressed in feet per second per second) by the square of the number of seconds through which the body travels and we ascertain the number of feet passed over by the body.

The last equation of all provides us with a rule for calculating the velocity of a body after it has travelled over a certain distance under the influence of a uniform acceleration; it tells us that the square of the body's velocity is equal to twice the product of the acceleration (in feet per second per second) under which it moves into the distance in feet it has travelled over.

Easy examples on the application of these rules will be found in the questions at the end of the chapter and in solving them the student must carefully bear in mind that all the formulæ given in the chapter are equally applicable if the centimetre and second, or other units, are adopted.

Chief Points of Chapter III.

The Solar Day is the interval of time between the sun's highest position on any one day to its corresponding position on the next succeeding day.

The Mean Solar Day is the quotient obtained by dividing the length of the year measured by the sun by the number of days in the year. Or, it is the average of the lengths of all the solar days in the year.

A Star or Sidereal Day is of constant length. It is the interval of time between two succeeding southings of a star at any season of the year.

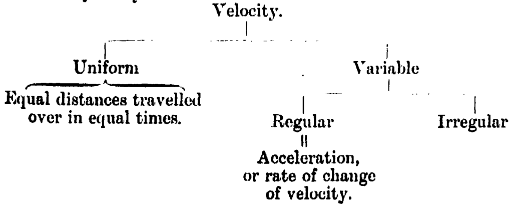
The Period of Rotation of the Earth.—As the apparent motions of the stars across the sky are produced by the rotation of the earth, the exact time of rotation is determined by ascertaining the length of the sidereal day.

Units of Time.—In physical measurements the unit of time adopted is the mean solar second, that is, it is founded on the average time required by the earth to make one complete rotation on its axis relatively to the sun considered as a fixed point of reference.

The Velocity of a body is the rate at which it moves. Velocity may be either *uniform* or *variable*. When a body's velocity is uniform it moves over the same distance in every second. When the velocity is variable unequal distances are moved over in equal times.

$$\text{Uniform velocity} = \frac{\text{number of units of length travelled}}{\text{number of units of time taken}}.$$

Acceleration is the rate of change of velocity. In measuring uniform acceleration we must know what addition to or subtraction from the velocity of the moving body there has been during each second of its journey.


Velocity.
Uniform
Equal distances travelled over in equal times.
Variable
Regular
Acceleration, or rate of change of velocity.
Irregular

The Unit of Acceleration is an increase of unit velocity in a unit of time. It is generally taken as equal to an increase of velocity of one foot per second per second.

The change of velocity v in t seconds in a body moving from rest with an acceleration of a units is equal to at feet per second. If a uniform acceleration is applied to a body already moving with a uniform velocity u, then $v = u + at$.

Other Important Formulæ.—

$$s = \frac{1}{2}at^2$$
$$v^2 = 2as$$

QUESTIONS ON CHAPTER III.

1. Give a definition of *velocity*, and state how uniform linear velocity is measured.

2. What is the difference between a *solar day* and a *mean solar day?*

3. Define a sidereal day and explain how it differs from a solar day.

4. What is the common unit of time? How is it related to the period of the earth's rotation?

5. Give a brief description of some mechanism in common use for measuring time.

6. What is the difference between *uniform* and *variable* velocities? How are uniform linear velocities measured?

7. Explain how the average velocity of a body moving with a variable velocity is estimated.

8. The radius of a circle is half a mile; a horse runs round the circumference at the rate of five miles an hour? what is his velocity in feet per second? (The result should be carried to the second decimal place.)

9. Define acceleration and give an example of a body moving with an accelerated velocity.

10. How is uniform acceleration measured? Explain the difference between a positive and negative acceleration, giving an example of each.

11. What is the unit of acceleration? A body moves from rest with a uniform acceleration of 15 feet per second per second; what is its velocity after 12 seconds?

12. The velocity of a body moving from rest has been uniformly accelerated; if its velocity after 7 seconds be 343 feet per second what has been its acceleration?

13. A train starting out of a station is made to move with a uniform acceleration of 25 kilometres per hour per hour; what is its velocity in kilometres per hour after it has gone a distance of 16 kilometres?

14. How far will a ball go from rest in three seconds if in every second it increases its velocity by a velocity of 9 feet per second?

15. A train which is uniformly accelerated starts from rest, and at the end of 3 seconds has a velocity with which it would travel through 1 mile in the next 5 minutes; find the acceleration.

16. A body moves from rest with a uniform acceleration of 100

inches per second per second; calculate its velocity after it has moved over 20 feet.

17. Explain the following statements:—A body has a velocity whose measure is 30; another body an acceleration whose measure is 2; in each case with foot second units. If allowed to move freely, how far would the first body travel in half a minute, and in what time would the second body, starting from rest, acquire a velocity of 15 miles an hour?

18. A particle which started from rest and was uniformly accelerated, had a velocity of 20 feet per second when it had described 60 feet in a straight line. What was its acceleration and how long had it been in motion?

19. Two bodies, whose velocities are accelerated in every second by 3 and 5 feet a second respectively, begin to move towards each other at the same instant, at first they are a mile apart; after how many seconds will they meet?

20. A man runs uniformly round a circular track at the rate of 12 miles an hour; if he pass the same point four times in 5 minutes, what is the length of the track?

21. Two stations on a railway are 35 miles apart. A train starts from the first and travels uniformly at the rate of 40 miles per hour towards the second; another train starts simultaneously from the second station and travels at the rate of 30 miles per hour in the opposite direction. At what time after starting will they meet and at what distance from the two stations respectively?

22. A man walking at the rate of 4 miles an hour along a railway line meets a train, 88 yards long, travelling at the rate of 26 miles an hour. How many seconds elapse while the train is passing him?

23. A line is drawn upon the floor of a railway carriage from door to door. When the carriage is at rest a ball is dropped from the roof and falls upon this line. What difference would be observed:—

(*a*) If the train is moving when the ball is dropped.

(*b*) If the train starts when the ball is half way down.

(*c*) If the ball is dropped when the train is in motion, but the train stops suddenly when it is half way down?

CHAPTER IV

MASS AND ITS MEASUREMENT

The mass of any body is the quantity of matter it contains In our country the standard or unit of mass is the quantity of matter contained in a lump of platinum of a certain size which is kept at the Exchequer Chambers. This amount of matter is called the **imperial standard pound avoirdupois,** and we speak of the mass of any other body as being a certain number of times more or less than the standard pound, that is, containing so many more times as much (or as little) matter as that contained in the imperial standard pound. Unfortunately, this is not a universal standard ; in France they have a standard of their own. It is kept at Paris, and is called a **kilogram,** and the system of masses founded upon it is used in all scientific work throughout the world.

Units of Mass in the British System.—As the reader knows very well, the unit of mass, the pound avoirdupois, to which attention has just been called, is not the only standard used in this country in speaking of the mass of a given body ; multiples and submultiples of it, as shown in the following familiar table, are also adopted. A comparison of this with the corresponding beautifully simple table of the metric system will do more than any words to show how cumbersome and inconvenient the British system is :—

16 drams	make	1 ounce (1 oz.).
16 ounces	,,	1 pound (1 lb.).
14 pounds	,,	1 stone (1 st.).
2 stones or 28 pounds	,,	1 quarter (1 qr.).
4 quarters or 8 stones	,,	1 hundredweight (1 cwt.).
20 hundredweights ...	,,	1 ton.

But this ungainly system of units does not exhaust the difficulties of the British system of estimating masses, for if the mass of a body composed of one of the precious metals, platinum, gold, or silver, or of some precious stone, is to be expressed another system known as "Troy Weight," is used. Moreover, the ounce in the last-named system is differently sub divided again if the masses of drugs are being dealt with.

Units of Mass in the Metric System.—In the metric system a name is given *to the mass of pure water which will exactly fill a cubic centimetre at a temperature of 4° C.* It is called a **gram**. The same prefixes are used to express fractions and multiples of a gram as have been used in the case of the metre and litre. The kilogram is one thousand times greater than one gram, and is the unit in use for ordinary purposes. Other multiples are in common use, as the following table shows :—

10 milligrams = 1 centigram.
10 centigrams = 1 decigram.
10 decigrams = **1 gram.**
10 grams = 1 dekagram.
10 dekagrams = 1 hectogram.
10 hectograms = 1 kilogram.
10 kilograms = 1 myriagram.
10 myriagrams = 1 quintal.
10 quintals = 1 millier.

Relation between British and Metric Units of Mass :—

Metric to British.

1 decigram = 1·5432 grains.
1 gram = 15·4323 grains.
1 dekagram = ·3215 oz. Tr.
1 hectogram = 3·5274 ozs. Av.
1 kilogram = 2·2046 lbs. Av.

British to Metric.

1 grain = ·648 decigram.
1 oz. Tr. = 3·1103 dekagrams.
1 oz. Av. = 2·835 dekagrams.
1 lb. Av. = ·4563 kilogram.

Gravitation.—It is a matter of everyday observation that masses free to move gravitate, as we say, towards the earth ; or as it is sometimes expressed, the earth attracts all bodies to itself. Experiments and observations made by Newton led him to the conclusion that it was the rule of nature for every body to attract every other body, and that this force of attraction is proportional to the body's mass, a large mass exerting a greater force of attraction than a small mass. But the farther these bodies are apart the less will be the attraction between them, though it is not less in the proportion of this distance, but in that of the square of the distance. This diminution of a force according to the inverse proportion of the square of the distance applies to so many cases that it ought to be clearly understood before going further. To give an example: two bodies of equal mass are one foot away from one another and attract each other with a certain force, call it a unit force. One body is now moved until its distance is two feet away from the second body, what will be the force of attraction between them? The square of 2 is $2 \times 2 = 4$ and the inverse of 4 is $\frac{1}{4}$, therefore the force of attraction is one quarter of the unit force. In the same way, if the bodies were three feet apart, the force of attraction would be $\frac{1}{9}$ of the unit force. Putting Newton's law together it stands thus:—**Every body in nature attracts every other body with a force directly proportional to the product of their masses and inversely proportional to the square of the distance between them; and the direction of the force is in the line joining the centres of the bodies.**

Returning to the case of the falling body, think of a cricket ball on the top of a house. The earth attracts the ball, and, by Newton's law, the ball attracts the earth. The ball, if free to move, falls to the earth; to be correct, however, we must think of the ball and the earth moving to meet one another along the line joining their centres. But the ball moves as much farther than the earth as the earth's mass is greater than the ball's; and for practical purposes this is the same as saying that only the ball moves and that the earth remains still. Were our methods of measurement sufficiently refined we should, of course, be able to measure the small amount of the earth's movement.

This force of attraction between all material bodies is called the force of gravity, and we must point out that this is only a name. Calling this force "gravity," and the rule according to which it acts the "law of gravity," does not teach us anything about the nature of the force itself. There is, however, no doubt about the reality of the attraction which bodies exert upon one another. So long ago as 1798 Cavendish measured the force of attraction between two lead balls, and his experiments have been repeated by a number of other investigators.

The weight of a body is the force with which it tends to move towards the earth. We see at once the difference there is between mass and weight. One is matter, the other is a force. When we speak of a mass of one pound we have seen that it means nothing more than the amount of matter in a certain lump of iron or other material; but when we speak of the weight of one pound we mean the force with which the mass of a pound tends to move towards the earth. The student will do well to clearly understand this distinction before proceeding.

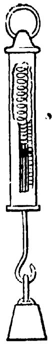


FIG. 20.—A Spring Balance.

EXPT. 46.—Procure a spring balance (Fig. 20) and attach it to a rigid support. Notice that to bring the index down you must exert a pull on the bottom hook. The greater the pull the farther down the scale does the index move.

EXPT. 47.—Attach any convenient mass to the hook of the balance, and notice that the index moves a certain distance down the scale as in the last experiment. The pull in this case measures the earth's attraction for, or the *weight* of, the mass at the place of observation.

EXPT. 48.—Repeat the last experiment, using different masses, and notice that the weight of the masses, as measured by the pull on the spring, varies in the different cases. Whenever the pull on the spring is the same, we know the masses are equal, and consequently the equality of masses can be tested by the equality of their weights.

The Weight of a given Mass varies from place to place.—Bearing the definition of weight in mind, it will be clear from Newton's law of gravitation that since a mass is farther away from the earth (which acts exactly as if its whole mass were collected at its centre) when it is on the top of a mountain than when at the sea-level, the weight of this mass ought to be more at the sea-level, for it is there nearer the centre than at the mountain top. This is found to be the case, but to actually demonstrate the difference in weight we must measure the weight by a spring balance as in the last experiments.

Similarly, because the earth is not a perfect sphere, but is flattened at the poles, points at the surface of the earth in the region of the tropics are at a greater distance from the centre than points in the neighbourhood of the poles. Consequently the weight of a mass situated on the earth in the tropics should be less than the weight it would have if it were moved into the polar regions. This has been found to be the case ; a body weighing 191 ounces when near the equator will weigh 192 ounces if moved near to the poles. These facts can be illustrated by a simple experiment.

EXPT. 49. Hang a piece of iron from a spring balance and notice the weight indicated. The iron has a certain mass, or consists of a certain quantity of matter. Bring a strong magnet under the iron, and again notice the indication of the pointer of the balance. Evidently the quantity of matter, or mass, of the iron has not changed during the experiment, but the attraction of the magnet causes the apparent weight to increase.

Masses are determined by means of the Balance.—Though the reader may have seen the spring balance in use among hawkers in the country, it is not a very common practice to use it for estimating masses. The instrument generally in use for this purpose is the **balance** or **pair of scales**. This simple apparatus is best understood with the help of a few simple experiments :—

EXPT. 50.—Balance a light stiff lath upon an edge of a triangular block, or better, make a hole through a point above its centre so that the lath will turn easily upon a stout nail fixed in a wall or blackboard. Hang a mass by means of a piece of thread upon the lath at any convenient distance on one side of the pivot or fulcrum, and balance it with a mass of the same amount on the other side. *The distance of the masses from the fulcrum will be found the same in each case.*

EXPT. 51.—Using the same lath supported on the nail at its middle point, hang over it, at equal distances from the support, two pans (which you can make out of pill boxes and cotton), one on each side of the support. Put a mass of 20 grams in one pan and ascertain how many must be put into the other in order that the lath may remain horizontal, or as we say, *be* *in equilibrium.* It is found that *when the pans are at equal distances from the supports there is equilibrium when the masses are equal.*

Expt. 52.—Repeat the last experiment, but while keeping the pan with the 20 grams in it in the same position, shift the other pan and find how many grams must be put into it to bring about equilibrium. Do this with the movable pan in a variety of positions and show that when the lath is balanced,

$$\begin{matrix}\text{Mass on} \\ \text{one side}\end{matrix} \times \begin{matrix}\text{Distance} \\ \text{from fulcrum}\end{matrix} = \begin{matrix}\text{Mass on} \\ \text{other side}\end{matrix} \times \begin{matrix}\text{Distance} \\ \text{from fulcrum.}\end{matrix}$$

Moments.—Another way of expressing the condition of things when there is equilibrium in Expt. 52 is to say that *the turning effects of the forces, about the point of suspension, acting on the lath at the points where the pans are supported, are equal.* The turning effect of a force is termed the *moment* of the force.

Referring to the condition of things in our experiment we can represent it very simply by a diagram (Fig. 21), where F is the point of support of the lath, and M_1 is a larger mass at a distance AF in equilibrium with a smaller mass M_2, at a greater distance FB.

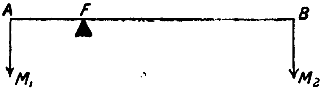


Fig. 21.—To Illustrate Moments of Forces.

The force at A is, as we know, the weight of the mass M_1 acting vertically downwards; and the force at B is that of the weight of the mass M_2 acting in the same direction. The moment of the force acting vertically downwards at A is a product like that obtained in Expt. 52, viz. the product of the force equal to the weight of M_1 and the distance AF, which as the diagram shows is measured at right angles to the direction in which the force acts. Similarly the moment of the force equal to the weight of the mass M_2, about the point A, is equal to the product of this force and the vertical distance BF.

This is a rule of universal application for taking moments which will have to be used several times later on, and must be very carefully learnt:—

The moment of a force about any point is the product

obtained by multiplying this force by the vertical distance between the point and the line of action of the force.

But it must always be remembered that the force and the vertical distance between its line of action and the point about which moments are taken must be expressed in suitable units.

The Balance.—The balance is really another form of the supported lath in Expt. 50. All the parts are very carefully made, and every means is taken to have very delicate supports and accurate adjustments. Fig. 22 shows a simple form of balance which is very suitable for easy experiments, such as are described in this book, which can all be accurately done by its means. Instead of the wooden lath in the experiment we have a brass beam supported at its middle line by a knife-edge of hard steel, which, when the balance is in use, rests on a true surface of similar steel. The hooks to which the pans are attached are similarly provided with a V-shaped depression of hard steel, which also, when the balance is in use rests upon knife-edges on the upper parts of the beam. To the middle of the beam is attached a pointer, whose end moves over an ivory scale fixed at the bottom of the upright which carries the beam. When not in use the beam and hooks are lifted off the knife-edges by moving the handle C.

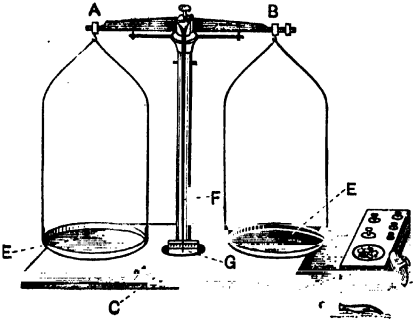


FIG. 22.—The Students Balance.

Fig. 23 shows a very delicate form of balance used by chemists. The principle of its action is precisely the same as in simpler kinds, but, to ensure greater accuracy, the knife-edges and the supports on which they rest are made of *agate*, a very hard mineral, instead of being constructed of steel.

EXPT. 53.—Uncover the balance and identify the different parts by reference to Fig. 22. Raise the beam, AB, of the balance, off the supports by turning the handle C. Notice whether the pointer F swings equally on both sides of the middle of the scale G: if it does the balance is ready for use; but if not, let down the beam and turn the small screw at B, then try again. Repeat this adjustment until the swings to right and left are equal.

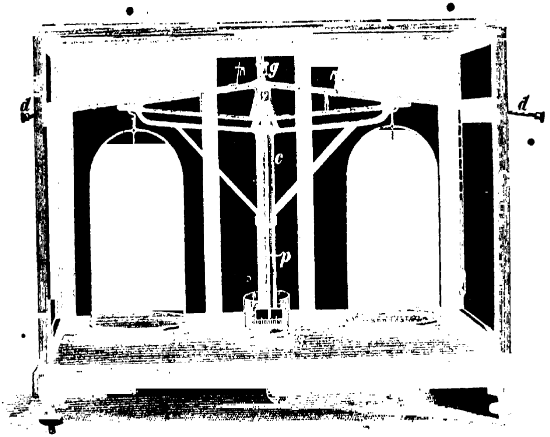


FIG. 23.—A Chemical Balance.

EXPT. 54.—Put some shot in a watch-glass, and place it in the *left-hand pan*, E, of the balance. Examine your box of weights, and, selecting a weight which you estimate to be about the same as the shot, take hold of it *with the forceps* and place it in the *right-hand pan* (the weights should always be placed in this pan). Now slightly raise the beam to see whether the estimated weight is nearly equal to the weight of the shot. If the weight appears a little below what is wanted, pick up, with the forceps, the next heaviest weight in the box, and try it in the pan with the other. If the two together are too heavy, take out the smaller weight and put in the one below it, and so on, adding one weight after the other, *without missing any*, until you find the correct weight. When you have completed the weighing, write down and add up the weights that are missing from their places, and check the figures as you put each weight back in its place.

EXPT. 55.—To become familiar with the process of weighing, find the mass of half-a-crown, a shilling, and other suitable bodies.

EXPT. 56.—Find the mass in grams, &c., of an ounce weight.

The Weight of a Body is not determined by an ordinary Balance.—The spring balance, and not a pair of scales, must be used to determine the weight of a mass at any place, because by means of the latter apparatus all we do is to make a comparison between the weight of an unknown and that of a known mass. Though the *weight* of the known mass (say the 50 gram weight out of the box) varies from place to place, its mass remains constant, and what we determine when we effect a weighing with the balance is the *mass* of the body experimented upon. If we could attach the 50 gram weight to a very sensitive spring balance and carry it from the equator to the pole, we should find, as we have already seen, that its weight as recorded by the spring balance would vary continuously. Hence we estimate the equality of masses by the equality of their weights; the weight of the body whose mass is to be determined must, when the balance is in equilibrium, be equal to that of the known mass, and consequently we can argue that their masses are equal. **By the balance we measure the mass of a body, whereas a spring balance enables us to determine its weight at a given place.**

When a Balance is in equilibrium the moments, about the point of suspension, of the forces equal to the weight of the masses are equal. When the balance is true the arms are of equal lengths, and consequently the weights are equal also.

CHIEF POINTS OF CHAPTER IV.

The **Mass** of any body is the quantity of matter it contains.

The **Imperial Standard Pound Avoirdupois** is the amount of matter contained in a lump of platinum of a certain size which is kept at the Exchequer Chambers.

The metric standard of mass is the **Gram**; but in everyday use the kilogram, which is equal to a thousand grams, is employed.

A **Gram** is the mass of pure water which will exactly fill a cubic centimetre at a temperature of 4° C.

The same prefixes are used to designate the multiples and sub-multiples of the gram as are used in the case of the metre.

Law of Gravitation.—Every body in nature attracts every other body with a force directly proportional to the product of their masses and inversely proportional to the square of the distance between them; and the direction of the force is in the line joining the centres of the bodies.

The **Weight** of a body is the force with which it tends to move towards the earth.

The Weight of a given mass varies from place to place.—Its weight at the sea-level is more than on the top of a mountain. Its weight is also greater when it is in the neighbourhood of the poles than when at the equator.

Masses are determined by means of the Balance.—It is found that when the pans of the balance are at equal distances from the supports, there is equilibrium when the masses are equal.

The **Moment of a Force** about any point is the product obtained by multiplying this force by the vertical distance between the point and the line of action of the force.

When a Balance is in equilibrium the moments, about the point of suspension, of the forces equal to the weights of the masses are equal. When the balance is true the arms are of equal lengths and consequently the weights are equal also.

Questions on Chapter IV.

1. Define the *mass* and *weight* of a material body, carefully distinguishing between the terms.

2. Give the British and metric measures of mass.

3. State Newton's law of gravitation, explaining clearly the rule of inverse squares.

4. What is a spring balance and what can be measured with it?

5. Explain fully what different readings, if any, are given by a spring balance under the following circumstances and account for them as far as you can.

(*a*) When a spring balance carrying a given mass is taken from the equator to the north pole.

(*b*) When the same balance is carried down a deep mine.

(*c*) When the balance is taken to the summit of a high mountain.

6. Describe an instrument for determining the mass of a material body and explain how it is used.

7. What do you understand by the *moment of a force*? Explain your definition by an example.

8. What is the condition of things, from the point of view of moments, when a balance with both pans loaded is in equilibrium.

9. What particulars about a given mass could you determine (*a*) by using a spring balance, (*b*) by means of a pair of scales?

10. Multiply 10·4 square centimetres by 15·5 decimetres, and state the result both in cubic centimetres and in litres.

If the volume in question were filled with water at 4° C., what would the weight of the water be?

11. A body, of which the volume is 1·5 litres, has a mass of 25 kilograms. What is the mass in grams of 1 cubic centimetre of the body.

What is the volume at 4° C. of 1,712 grams of water?

CHAPTER V

MATTER IN RELATION TO MOTION

Equality of Masses.—We have hitherto considered motion and mass separately, but in this chapter their relations to one another will be described. It has been shown that equality of mass can be tested by weighing. The balance thus furnishes us with a convenient practical method of comparing masses, but it does not give a fundamental conception of what mass means. To obtain a clear idea of the subject, consider first of all that we are dealing with two variable quantities, namely, mass, or quantity of matter, and motion. Suppose two bodies moving in opposite directions with equal velocities to collide with one another and stick together. If the two bodies stopped dead after the impact we could conclude that their masses were equal, and that each exactly destroyed the motion of the other; but if the combined bodies moved after the collision, the masses could evidently not have been equal. With this in mind, it will readily be conceded that the following definition of equality of mass holds good:—

Two masses are equal, if when they are made to impinge on one another in opposite directions with equal speeds and stick together, they come to rest.[1]

The effect of the impact of two moving bodies thus depends upon the masses of the bodies and the velocities before the collision. If both the velocities and masses are equal, the bodies come to rest; if the velocities are equal, but the masses are unequal, the greater mass predominates after collision, and if the masses are equal while the velocities are unequal the greater velocity will predominate.

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[1] This definition and one or two others in this chapter are given by Prof. W. M. Hicks in his inspiring work on *Elementary Dynamics of Particles and Solids.*

Momentum.—The preceding paragraph has introduced the student to the idea of a condition involving both motion and mass. This condition is known as momentum, and is defined as follows :—

The momentum of a body is the quantity of motion it has, and is equal to the product of its mass and its velocity.

Expressed as an equation we have

Momentum = mass × velocity,

or if momentum is represented by M, mass by m, and velocity by v, all expressed in corresponding units, we can write

$$\mathbf{M = mv.}$$

The unit of momentum is consequently that of a unit of mass moving with a unit of velocity, or if the unit mass be that of the imperial standard pound, **the unit of momentum is the quantity of motion in a mass of one pound moving with a velocity of one foot per second.** The meaning of momentum will be better grasped after a concrete example.

Suppose a shot fired from a cannon, the momentum generated in both the cannon and the shot will be the same ; but since the mass of the cannon is immensely greater than that of the shot, it will be evident that the velocity of the shot must be correspondingly greater than that of the cannon in order that the product of the two quantities may be the same. This we know is the case, the velocity of the "kick" or "recoil" of the cannon is very much less than the velocity with which the shot is sent on its journey.

The motions of the shot and of the cannon are in opposite directions, and it is usual to distinguish the two directions by the signs + and − . If the mass of the cannon is represented by M, and the mass of the shot by m, while the velocities of the two are v and V respectively, then the momentum after the explosion is expressed by the equation

$$Mv = -mV.$$

EXAMPLE.—A cannon weighing 10 tons fires a shot weighing 28 lbs. The shot leaves the gun with a velocity of 200 feet per second. With what velocity would the cannon recoil if it were free to move?

$$\begin{aligned} \text{Weight of cannon} &= 10 \text{ tons} = 22{,}400 \text{ lbs.} \\ \text{Velocity of recoil} &= v = \,? \\ \text{Weight of shot} &= 28 \text{ lbs.} \\ \text{Velocity of shot} &= 200 \text{ feet per second.} \end{aligned}$$

We know that

$$\begin{aligned} \text{Momentum of cannon} &= \text{momentum of shot} \\ \text{Or, } 22{,}400 \times v &= 28 \times 200 \\ \text{Therefore } v &= 5{,}600 \div 22{,}400 \\ &= \tfrac{1}{4} \text{ foot per second.} \end{aligned}$$

The Conservation of Momentum.—It is convenient in many problems to consider momentum from another point of view, namely, that of the conservation of momentum, by which is meant that though momentum can be transferred from one body to another, none of it is lost by the transference. For instance, if a number of inelastic bodies collide with one another, and then move on together, the total momentum of the moving mass is equal to the sum of the momenta of the separate bodies. In other words, the total momentum of several moving masses remains unaltered by impact. Expressed as an equation, in which M and M are different masses moving with velocities v and v', while the velocity after impact is V, we have

$$(\mathbf{M} + \mathbf{M'})\,\mathbf{V} = \mathbf{Mv} + \mathbf{M'v'}.$$

The use of this equation will be best illustrated by a few arithmetical examples.

EXAMPLE.—Two inelastic balls, having masses of 4 lbs. and 7 lbs., and moving in opposite directions with velocities of 8 and 10 feet per second respectively, collide with one another. What is the velocity of the joint mass after impact?

As the velocities are in opposite directions, one of them must be distinguished by the sign $-$. The equation thus becomes

$$\begin{aligned} (M+M')\,V &= Mv - M'v' \\ (4+7)\,V &= (4 \times 8) - (7 \times 10) \\ 11\,V &= 32 - 70 \\ &= -38 \\ \text{Therefore } V &= -38 \div 11 \\ &= -3\tfrac{5}{11} \text{ feet per second} \end{aligned}$$

The mass will therefore move with a velocity of $3\frac{5}{11}$ feet per second in the direction in which the 7 lbs. was moving.

EXAMPLE.—Two masses of 6 and 10 lbs., moving with velocities of 4 and 5 feet per second respectively, collide with one another. Find the final velocity, (*a*) if the bodies are moving in the same direction; (*b*) if they move in opposite directions.

The same equation is used as before, the only difference being that in the first case the velocities are both taken as +, whereas in the second case one is taken as −. We therefore proceed as follows:—

Case (*a*)—

$$\begin{aligned} (6+10)\,V &= (6\times4)+(10\times5) \\ 16\,V &= 24+50=74 \\ \text{Hence } V &= 74\div16 \\ &= 10\tfrac{5}{8} \text{ feet per second.} \end{aligned}$$

Case (*b*)—

$$\begin{aligned} (6+10)\,V &= (6\times4)-(10\times5) \\ 16\,V &= -26 \\ \text{Hence } V &= -26\div16 \\ &= -1\tfrac{5}{8} \text{ feet per second.} \end{aligned}$$

In this case the direction of motion is that of the mass of 5 lbs.

EXAMPLE.—A bullet weighing 2 oz., and moving with a velocity of 1,200 feet per second, is fired into a block of wood weighing 2 cwt., and carries the wood onwards with it. With what velocity does the wood move?

$$2 \text{ cwt.} = 3584 \text{ oz.}$$

$$\begin{aligned} \text{Momentum after event} &= \text{momentum before event} \\ (2+3{,}584)\,V &= (2\times1{,}200)+(3{,}584\times0) \\ 3{,}586\,V &= 2{,}400+0 \\ \text{Therefore } V &= 2{,}400\div3{,}586 \\ &= 0\cdot67 \text{ feet per second.} \end{aligned}$$

Force.—Suppose a body to possess a certain momentum, then for the momentum to change or tend to change, something must act upon the body, and that something is termed *force*. In the words of Professor Hicks: **When a gradual change of momentum is either produced or tends to be produced in a body, that body is acted on by force.**

It must be clearly understood that by thus defining force we do not get to know anything more about it. Nobody can tell what force is. All we can know are the effects produced by a something we call force.

Unit of Force.—A change of momentum is produced by force; the rate at which the momentum changes may therefore be used as a measure of force. The unit of force can be defined in several ways.

A unit of force acting for the unit of time is able to produce a unit of velocity in a unit of mass.

Or, a unit of force produces a unit of acceleration in a unit of mass. But since the product of a mass and its velocity is spoken of as the momentum of the body, we can measure force by the momentum it generates, *the unit force giving rise to the unit of momentum in the unit of time.* **Equal forces are, therefore, those which produce equal momenta in equal times.**

The momentum generated by a force of two units is twice as great as that produced by one unit; and, further, a force of one unit acting for two seconds will produce twice the momentum which it would do if it only acted for one second. This is why it is necessary in defining the unit of force to introduce the words "acting for the unit of time."

Acceleration produced by a Force.—The momentum of any particular body is determined by the body's mass and velocity. As the mass of the body may be regarded as constant, change of momentum can only be produced by changing the velocity. But rate of change of velocity is acceleration, hence when the acceleration of a body is altered, the momentum is altered, and an alteration of momentum signifies, as has been explained, that the body is being acted upon by a force. If the acceleration is uniform, the body must be acted upon by a uniform force.

Hence we come to the very important fact that **the number of units of force in any force is equal to the product of the number of units of mass in any body on which it may act and the number of units of acceleration produced in that mass by the force in question.**

The relation between force, mass, and acceleration may be expressed algebraically as follows:—Let F represents the number of units of force in a given force, m the number of units of mass on which it acts producing a units of acceleration, then our definition can be written,

$$\mathbf{F = m \times a},$$

from which equation the third quantity can be obtained whenever we know the other two :—

$$\begin{array}{c} \text{Number of} \\ \text{units of force} \end{array} = \begin{array}{c} \text{Number of} \\ \text{units of mass} \end{array} \times \begin{array}{c} \text{Number of units} \\ \text{of acceleration.} \end{array}$$

$$F = ma \quad \text{........................(1)}$$

$$\therefore \quad a = \frac{F}{m} \quad \text{........................(2)}$$

$$\therefore \quad m = \frac{F}{a} \quad \text{........................(3)}$$

The second equation can be expressed in words by saying that the number of units of acceleration produced in the velocity of a moving body is equal to the number of units of force acting upon it, divided by the number of units of mass on which it acts.

Similarly, the third expression means that the number of units of mass in a moving body can be calculated by dividing the number of units of force acting upon it by the number of units of acceleration produced in it.

The second equation tells us, moreover, that if the acceleration produced in a moving body remains the same, or is uniform, that the value of the force, or the number of units of force it contains, must be the same throughout, or what is the same thing, the force is uniform.

The most important facts with reference to motion and force have been described in the foregoing parts of this chapter from the point of view of momentum. At this point the first law of motion may be usefully introduced.

Newton's First Law of Motion.—Every body will continue in its state of rest or of uniform motion in a straight line, except in so far as it is compelled by impressed force to change that state.

This statement is called a *law* of motion, and it will be well to make clear at once that a natural law is only an expression of what has been found to always be the rule ; it is merely setting down the result of experience ; and the idea of the word in its legal sense must be carefully excluded from the mind. Indeed, in nearly every case, it is better to substitute, at all events in one's mind, the word *rule.* It is only a statement that certain things always seem to take place ; it tells us nothing about why they do so ; nor is the idea of compulsion included at all, for oftentimes so-called "laws" have been formulated which have turned out to be wrong expressions of the order of nature, and evidently there could be no sort of compulsion about what was wrong, and was seen afterwards to be contrary to the general rule.

This law, which Newton first stated as being always followed by bodies in nature, means, first, that if a body is at rest, it will remain still until there is some reason for its moving—until some outside influence, which is called a *force*, acts upon it. In fact the law really supplies us with a definition of force. Nobody finds any difficulty in understanding the rule so far. But when we come to consider the second part of the law there is more difficulty in grasping it—every body will continue in a state of uniform motion in a straight line, &c. An example is afforded by a ball in moving uniformly along ice. We know that after a time the ball comes to rest and therefore that it does not continue in a state of uniform motion. But we know that it moves for a longer time on ice than it would do on a road. The ice is smoother than the road, and there seems to be a connection between the roughness or smoothness and the length of time during which the ball moves. If we imagine smoother and smoother ice, the ball will move for a longer and longer time, and we conclude that if both the ball and the ice were perfectly smooth, there is no reason why the ball should ever stop. The roughness or friction is then, in our example, the "impressed force" which causes the ball to change its state of uniform motion for one of rest. If we could have a body in a state of uniform motion outside the influence of what Newton has called "impressed forces" it would afford us an example of *perpetual motion*. But because we cannot eliminate these impressed forces we cannot have perpetual motion.

The first law of motion implies the existence of force, which may consequently be defined as that which produces, or tends to produce, motion in matter; or alters, or tends to alter, the existing motion of matter.

The inability shown by a material body of itself to change its condition of rest or of uniform motion is called its inertia. Inertia may also be defined as the capacity of a body to possess momentum. We become aware of the inertia of matter most unpleasantly if we step out of a moving train on to the platform; while in the train we partook of its motion, on stepping out our feet are brought to rest suddenly, but our bodies, because of their inertia, continue to move with the velocity of the train, with the result that we fall forwards on to our faces. The difficulty of moving a heavy body, or giving it momentum, is also a familiar example of the same property.

The Attraction of Gravity at any place is an Example of a Uniform Force.—The general laws which have just been considered are well illustrated by the attraction of the earth upon a mass at its surface. We have learnt to call this force of attraction by the name of the body's weight. Let a body containing m units of mass be attracted to the earth with W units of force, or, as we may just as correctly say, let the weight of m units of mass be W units and call the acceleration which would be produced in it in the unit of time, if the mass moved towards the earth, g units; then by the equations given on p. 60 we can write

$$\mathbf{W = m\,g}$$

$$\text{and} \quad m = \frac{W}{g}$$

which tells us that **the weight of a body is equal to the number of units of mass it contains multiplied by the number of units of acceleration produced by gravity in one second in a body moving freely towards the earth.**

Value of "g."—It has already been seen that a body moving from rest with an acceleration of a units per second per second travels over a distance s represented by the equation: $s = \frac{1}{2}at^2$. Hence, in the case of a body moving towards the earth under the influence of the force due to gravity, which causes a regular acceleration of g units, the distance travelled over depends upon the value of "g." But the distance moved over by such a body moving from rest towards the earth in one second can be measured. It has been found to be 16 feet. All that has to be done, therefore, is to substitute this value in the equation above, after changing the general symbol for acceleration, a, into the particular acceleration, g, applicable to this case. We thus have to substitute 16 for s and 1 for t in the equation

$$\begin{array}{ll} s = \frac{1}{2}gt^2 & \text{which gives} \\ 16 = \frac{1}{2} \times g \times 1 & \text{from which} \\ g = 32 & \text{or} \end{array}$$

the acceleration due to gravity is equal to a velocity of 32 feet per second in every second.

Motion of Bodies falling from Rest.—The consideration of all questions concerning bodies falling from rest towards the earth is consequently only a particular case of the general rule which has already been learnt by the student. We have a constant force, that, viz., equal to the attractive force of the earth, causing a constant acceleration of 32 feet per second in every second, which is, as we have seen, the value of the acceleration due to gravity, g. All we have to do, therefore, is to substitute g for a in each of the equations on p. 40, and we get :—

$$\mathbf{v} = \mathbf{gt} \quad \ldots\ldots\ldots\ldots \quad (1)$$
$$\mathbf{s} = \tfrac{1}{2}\mathbf{gt}^2 \quad \ldots\ldots\ldots\ldots \quad (2)$$
$$\mathbf{v}^2 = 2\mathbf{gs} \quad \ldots\ldots\ldots\ldots \quad (3)$$

Velocity of a body falling from rest after "t" seconds.—The first of the above equations provides us with an expression by means of which, knowing the number of seconds, (t), for which a body has been moving freely towards the earth under the constant acceleration (g) due to the force of gravity, we can calculate the velocity with which it is moving at a given moment.

Example.—With what velocity is a body moving which starting from rest has travelled for 10 seconds?

Here we have to find v, knowing the value of both g and t.

$$v = gt$$
$$= 32 \times 10 = 320 \text{ feet per second.}$$

Distance travelled by a Body falling from Rest in "t" Seconds.—The second equation at once enables us to determine the distance through which a body, falling freely towards the earth, has moved after it has been travelling for a known number of seconds.

Example.—Through how many feet does such a body fall in 30 seconds?

$$\text{Here} \quad s = \tfrac{1}{2} gt^2$$
$$= \tfrac{1}{2} \times 32 \times (30)^2$$
$$= \tfrac{1}{2} \times 32 \times 900$$
$$= 14{,}400 \text{ feet.}$$

Velocity of a Body falling from Rest after moving through s Feet.—Equation (3) shows how the velocity of a body, which starts falling from rest, can be determined when the number of feet through which it has fallen is known :—

EXAMPLE.—What is the velocity of a stone which has fallen from rest for a distance of 100 feet?

$$\begin{aligned} \text{Here} \quad v^2 &= 2gs \\ &= 2 \times 32 \times 100 \\ &= 6400 \\ \therefore \quad v &= 80 \text{ feet per second.} \end{aligned}$$

Parallelogram of Forces.—A body can only move in one direction at any given moment; but yet it can be under the influence of, or be acted upon by, any number of forces. The question arises—How can the direction be determined in which a body, under the influence of several forces at the same moment, will move, if free to move?

Or, the question may be regarded in another way. How can we substitute a single force (called the **Resultant**) which will produce the same effect, for all the separate forces acting together upon a body at the same moment?

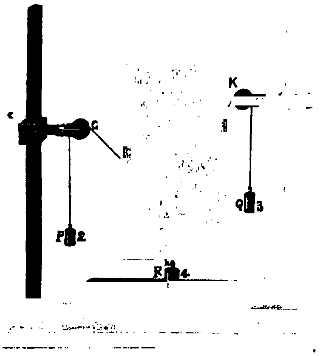

FIG. 24.—Experiment to illustrate the Parallelogram of Forces.

A few experiments will make the question and its answer quite clear as far as substituting a single force for two separate forces is concerned.

EXPT. 57.—Round two pulleys G and K, or very smooth pegs (Fig. 24), pass a fine thread to which two unequal weights are attached. To some convenient place on the thread tie a third weight as shown. Let the masses come to rest. It will be found that a parallelogram may be constructed, the sides and diagonal of which are nearly proportional to the weights used. They would be just proportional if the pulleys were quite smooth.

EXPT. 58.—Attach a scale of inches to the edge of a blackboard. Obtain two pieces of thin india-rubber cord twenty inches long, and fasten small loops of string to the two ends. Pin one of these loops to the board so that the upper end of the india-rubber cord coincides with the zero of the scale. Attach the upper end of the other india-rubber cord to any convenient point on the board. Bring the two lower ends together and hook on to them a weight (say of 100 grams). Measure off twenty inches from the upper end of each cord. The excess of length in each cord will be proportional to the tension of that cord. Complete the parallelogram with chalk, and show that the diagonal is vertical and is equal to the extension of the cord when it hangs vertically by the side of the scale with the weight attached.

These experiments lead to the rule which is always known as the *parallelogram of forces*. It is usually stated thus :—**If two forces acting at a point be represented in magnitude and direction by the adjacent sides of a parallelogram, the resultant of these two forces will be represented in magnitude and direction by that diagonal of the parallelogram which passes through this point.**

A line may be drawn to graphically represent a force, its length being made proportional to the force, and its direction showing the direction of the force. Forces can thus be represented, both in magnitude and direction, by lines. Let O represent a material body acted upon by two forces, represented both in amount and direction by the lines OB, OA. To find the **resultant** of these two forces as the single force which can replace them is called, both as regards its amount and direction, we complete the parallelogram OBRA and join OR, which will be the resultant required.

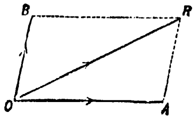

FIG. 25.—The Parallelogram of Forces.

Calculation of Resultant.—When the two forces whose resultant is required act at right angles to one another, the calculation is a simple application of a proposition in the first book of Euclid (I. 47). Under these circumstances the triangle ORA is right-angled, and Euclid proves that $(OA)^2 + (AR)^2 = (OR)^2$, and consequently $(OA)^2 + (OB)^2 = (OR)^2$, from which when OA and OB are known we can calculate OR.

When the directions of the two forces OB and OA are inclined to each other at an angle which is not a right angle, the calculation involves an elementary knowledge of trigonometry. This can be obviated, however, by the simple expedient of what is called the graphical method. This consists in drawing two lines inclined at the angle at which the directions of the forces are inclined, and making them of such length that they contain as many units of length as the force does units of force. The parallelogram is then completed by drawing AR and BA parallel respectively to OB and OA and joining the diagonal OR, whose direction will be that of the resultant, and whose length will be as many units as there are units of force in the resultant force.

Resolution of Forces.—A single force can be replaced by other forces which will together produce the same effect. Such a substitution is called *resolving* the force or a *resolution* of the force. The parts into which it is resolved are spoken of as *components.* When this has been done it is clear that we have made the original force become the resultant of certain other forces which have replaced it. Referring back to what has been said about the parallelogram of forces, it will be seen that any single force can have any two components in any directions we like; for by trying, the student will be able to make any straight line become the diagonal of any number of different parallelograms. The most convenient components into which a force can be resolved are those the directions of which are at right angles to each other. In this method of resolution, neither component has any part in the other.

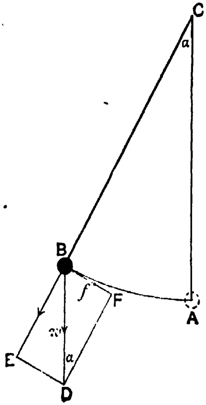

Fig. 26.—Resolution of a Force.

An interesting example of the resolution of a force into two components at right angles is afforded by a pendulum. Consider a pendulum at any point in its swing, as shown in Fig. 26. The pendulum-bob is pulled downwards in consequence of the attraction of gravity, and this vertical force is represented by the line BD. The force can, however, be resolved into two forces, one represented by *f* which pulls the pendulum in the direction BF, and the other represented by BE which has no part in moving the pendulum, and merely causes tension in the string BC.

CHIEF POINTS OF CHAPTER V.

Equal Masses.—If two bodies, moving with equal velocities in opposite directions, stop dead after colliding, we can conclude that their masses are equal. If the combined bodies moved after the collision their masses could evidently not have been equal.

Two masses are equal if when they are made to impinge on one another in opposite directions with equal speeds, and stick together, they come to rest.

The **momentum** of a body is the quantity of motion it has and is equal to the product of its mass and its velocity.

The **unit of momentum** is the quantity of motion in a mass of one pound moving with a velocity of one foot per second.

The **total momentum** of several moving masses remains unaltered by impact. This can be expressed by an equation (p. 57).

$$(m+m')V=mv+m'v'$$

Force.—When a gradual change of momentum is either produced or tends to be produced in a body, that body is acted on by force.

Unit of Force.—The unit of force can be defined in several ways:

1. A unit force acting for the unit of time is able to produce a unit of velocity in a unit of mass.

2. A unit of force produces a unit of acceleration in a unit of mass.

3. A unit of force gives rise to a unit of momentum in a unit of time.

A *poundal* is a force which produces an acceleration of one foot per second per second in the mass of a pound, and a *dyne* is the force which, acting upon a mass of one gram, produces an acceleration of one centimetre per second per second.

Equal Forces produce equal momenta in equal times.

The number of units of force in any force is equal to the product of the number of units of mass in any body on which it may act, and the number of units of acceleration produced in that mass by the force in question.

If F represents the number of units of force in a given force, m the number of units of mass on which it acts producing a units of acceleration, then we can write

$$F=m\times a$$

Newton's First Law of Motion.—Every body will continue in its state of rest or of uniform motion in a straight line, except in so far as it is compelled by impressed force to change that state.

Inertia.—The inability shown by a material body of itself to change its condition of rest or of uniform motion is called its inertia. Inertia may also be defined as the capacity of a body to possess momentum.

The Attraction of Gravity at any place is an example of a Uniform Force.—Let a body containing m units of mass be attracted to the earth with W units of force and call the acceleration which would be produced in it if the mass moved towards the earth g units. Then,

$$W = m\,g \quad \text{(p. 62)}$$
$$\text{and } m = \frac{W}{g}$$

The **Weight of a body** is equal to the number of units of mass it contains multiplied by the number of units of acceleration produced by gravity in a body moving freely towards the earth.

The **acceleration due to gravity** is equal to a velocity of 32 feet per second in every second, or of 981 cm. per second per second.

Motion of Bodies falling from Rest.—Let v = velocity with which such a body is falling, t the number of seconds through which it falls, and g the acceleration due to gravity. Then,

$$v = gt$$
$$s = \tfrac{1}{2}gt^2$$
$$v^2 = 2gs$$

Parallelogram of Forces.—If two forces acting at a point be represented in magnitude and direction by the adjacent sides of a parallelogram, the resultant of these two forces will be represented in magnitude and direction by that diagonal of the parallelogram which passes through this point.

Resolution of Forces.—A single force can be replaced by other forces which will together produce the same effect. Such a substitution is called *resolving* the force, and the parts into which it is resolved are called *components*.

QUESTIONS ON CHAPTER V.

1. Explain fully the circumstances which would justify us in declaring the masses of two bodies to be equal.

2. What is meant by momentum? Define the unit of momentum.

3. There are two bodies whose masses are in the ratio of 2 to 3 and their velocities in the ratio of 21 to 16. What is the ratio of their momenta? If their momenta are due to forces P and Q acting on the bodies respectively for equal times, what is the ratio of P to Q?

4. The masses of two bodies (P and Q) are in the ratio of 3 to 2; the former is moving at the rate of 7½ miles an hour, the latter at the rate of 200 yards a minute. Find the ratio of P's momentum to Q's momentum.

5. What is meant by Force? What are the results of the action of force upon a mass?

6. Define the *Unit of Force*. When are two forces said to be equal?

7. How is the number of units of force in any force estimated?

8. A force F acting on a mass of 5 lbs. increases its velocity in every second by 12 feet a second; a second force F_1 acting on a mass of 28 lbs. increases its velocity in every second by $7\frac{1}{2}$ feet a second. Find the ratio of F to F_1.

9. Enunciate Newton's First Law of Motion. Explain what you understand by a natural law.

10. Explain what is meant by the inertia of a material body. Give as many of the results of the possession of this property by a material body as you can.

11. How is the weight of a given body estimated? Why is the weight of a body different at the equator and at the poles?

12. How would you proceed to determine the value of the acceleration due to gravity?

13. Two stones are allowed to drop from the same height but the second stone 5 seconds after the first. How far apart will they be at the end of 2 seconds from the fall of the second stone?

14. Find the velocity of a ball thrown into the air 2 seconds after it has passed its highest point.

15. A body falls from a certain height and on reaching the ground has a velocity of 120 feet per second. What is the height?

16. What is the depth of a well if it takes 3 seconds for a stone to fall from the top to the bottom.

17. A body falls freely from rest. What is its velocity at the end of 1, 2, 3 seconds respectively?

18. A body falls from a tower 144 feet high. Find how long it takes to reach the ground and its velocity just before reaching it?

19. State the proposition known as the Parallelogram of Forces and give an experimental proof of it.

20. What is meant by the resultant of two forces acting at a point; and how can it be determined?

21. The horizontal and vertical components of a certain force are equal to the weights of 60 lbs. and 144 lbs. respectively. What is the magnitude of the force?

22. Forces of 5 and 12 units respectively act at a point. What are their directions when their resultant is greatest and when it is least?

Find the resultant when greatest, when least, and when the forces act at right angles to each other.

23. Draw two lines Ox, Oy at right angles to each other; two forces act at O; one of 7 units from x to O and one of 10 units from O to y; draw to any scale straight lines to represent these forces; draw, to the same scale, the straight line (OR) that represents their resultant; and find from the diagram the number of units of force in the resultant, and the number of degrees in the angle xOR.

24. Resolve a force of 56 lbs. into two others, so that the smaller

of the two shall be 33 lbs., and shall act in a direction at right angles to their resultant.

25. Two strings, the lengths of which are 12 inches and 16 inches, have their ends fastened to two points in the same horizontal line; their other extremities are fastened together and they are strained tight by a force equivalent to 5 lbs. at the knot, acting vertically; the angle between the strings is then found to be a right angle. Find the tensions of the strings.

26. The resultants of two forces, when they act in the same and opposite directions, are $5P+7Q$ and $P+Q$ respectively. Find the forces and show that, if each force is increased by a force P, the resultant of the two new forces acting at a point at right angles to one another is $5(P+Q)$.

27. Two lines AOB and COD are at right angles; and four forces act at O, one of 2 lbs. along OA, one of 3 lbs. along OC, one of 5 lbs. along OB, and one of 7 lbs. along OD. Find the value of the resultant and show, geometrically, how to find its direction.

28. A force 65P is resolved into two forces at right angles to one another. If one of them is 16P, what is the other?

29. Forces of 7 and 16 units have a resultant of 21 units; find the directions of the forces by a construction drawn to scale.

30. Draw an equilateral triangle ABC; a force of 10 units acts from A to B, and one of 15 units from A to C. Find their resultant by construction. Also find what their resultant would be if the force of 15 units acted from C to A.

31. A force equal to the weight of 20 lbs., acting vertically upwards, is resolved into two forces, one of which is horizontal and equal to the weight of 10 lbs. What is the magnitude and direction of the other component?

32. Draw a diagram, as well as you can, to scale, showing the resultant of two forces equal to the weights of 7 and 11 lbs. acting on a particle with an angle of 60° between them; and by measuring the resultant, find its numerical value.

33. Two forces, each equal to a weight of 10lbs., act in directions making an angle of 120°. Find their resultant.

34. A boat is towed by two ropes, including an angle of 60°. One rope is pulled with a force of 20 lbs., and the other with a force of 55 lbs. Find the resultant force which moves the boat.

35. A ball weighing 24 lbs. is kept at rest by two strings, one of which is horizontal and the other inclined to the vertical at an angle of 30°. Find the tensions of the two strings.

36. Four forces, A, B, C, D, of 1, 2, 3, and 4 lbs. respectively, act at a point. B acts at right angles to the direction of A; C to that of the resultant of A and B; D to that of the resultant of A, B and C. All the right angles are to be measured in the same angular direction. Draw, as accurately as you can, a figure showing how the resultant of all four forces may be found graphically; and show independently, by calculation, that its magnitude is $\sqrt{30}$.

37. The same force F is made to act continuously on two separate masses of 15. lbs and 75 lbs.; compare the accelerations produced.

38. Two equal masses of 6 lbs. each are connected by different

strings passing over the edge of a table with two other masses of 3 lbs. and 18 lbs. respectively. What are their accelerations and which will reach the edge of the table first if both are placed at a distance of 12 feet from that edge?

39. Two masses of 5 lbs. and 7 lbs. are connected by a light string which passes over a smooth pulley. Find their common acceleration and the velocity of either after they have been moving for 3 seconds. (The student must think here what is the moving force? and what is the mass on which it acts?)

40. A certain force F acting for 10 seconds produces in a mass of 17 lbs. a velocity of 12 feet per second; another force F′, acting for 10 minutes, produces in a mass of 17 cwt. a velocity of 12 yards per minute. Find the ratio of F to F′.

41. One force acting on a mass of 12 lbs. gives it a velocity of 5 feet per second, in one-third of the time that another force acting on a mass of 9 lbs. takes to give it a velocity of 8 feet per second. Compare the forces.

42. What force will set up in a mass of 5 lbs. a velocity of 80 feet per second in 5 seconds?

43. How many units of force (foot-pound-second system of units) must act on a mass of 16 lbs. for 4 seconds so as to give it a velocity of 35 yards per minute?

44. A certain force acting on a mass of 10 lbs. for 5 seconds, produces in it a velocity of 100 feet per second. Compare the force with the weight of 1 lb., and find the acceleration it would produce if it acted on a ton.

45. Equal forces act for the same time on two bodies A and B, the mass of the first being four times that of the second. What is the relation between the momenta generated by the forces?

46. A force moves a body 18 feet in 3 seconds, starting from a position of rest. Find the ratio of the force to the weight of the body.

47. A pressure equal to the weight of 3 lbs. acts on a mass of 8 lbs. starting from rest. Find the distance over which the mass will move in 4 seconds.

48. A horizontal force which would statically support 5 lbs., acted continuously for 3 seconds on a heavy body initially at rest on a smooth horizontal plane, and at the end of that time the body was moving with a velocity of 200 yards per minute. Determine (*a*) the acceleration and (*b*) the mass of the body.

49. A mass of 12 lbs. rests on a horizontal table. Find the pressure of the mass upon the table when the table is made to descend with a uniform acceleration of 5 feet per second per second.

50. A man whose weight is 160 lbs. is standing in a lift. What pressure will he cause on the bottom of the lift when it is (*a*) ascending, (*b*) descending with uniform acceleration $\frac{1}{8}g$?

51. Ten pounds hangs by a string from a point which ascends with an acceleration of 2 (feet per second per second); find the tension of the string.

52. A bullet, weight 1 oz., is fired from a rifle; if it has a velocity

of 1,520 feet per second on leaving the barrel, what will be its momentum at that moment?

53. Compare the momenta of two bodies, masses 4 tons and 4 lbs. respectively, the former having a velocity of 20 miles an hour and the latter 20 feet per second.

54. If a mass of 16 lbs. was moving at the rate of 100 yards a minute and a mass of 4 lbs. at the rate of 6 miles an hour, find the ratio of the momentum of the first to that of the second.

55. A bullet, mass $1\frac{1}{2}$ oz., moving horizontally with a velocity of 1,248 feet per second, strikes a cubical block of wood, mass 25 lbs., lying on a smooth horizontal plane, and remains imbedded in it; find the velocity with which the wood begins to move.

56. If a 500-lb. shot be fired by a 32-ton gun so that the shot leaves the gun with a velocity of 1,400 feet per second, find the velocity of the gun's recoil.

57. Just as a tramcar reaches a man standing by the tramway, it has a velocity of $8\frac{1}{2}$ feet per second, the man takes hold of and mounts the car. What change of velocity takes place?—the weights of the car and man being 1 ton and 10 stones respectively.

58. Two masses are moving with velocities of 12 feet per second and 40 yards per minute respectively; if their momenta be equal and the mass of one be 6 lbs., what is the mass of the other?

59. A mass of 8 lbs. moves from rest with an acceleration of 10 feet per second per second; what is its momentum (*a*) after 20 seconds, (*b*) after it has passed over a distance of 20 feet?

60. The same body is weighed by means of a spring balance in London and Rio de Janeiro; if the weighings show respectively 20 lbs. 1 oz. and 20 lbs. 2 oz., compare the values of *g* obtained.

61. The intensity of gravity on Jupiter is 2·5 times that on the earth; find the time it will take a body to fall 160 feet on Jupiter.

62. A force F acting on a mass of 10 lbs. increases its velocity in every second by 10·7 feet per second in a locality where the velocity gained by a falling body in a second of time is 32·1 feet per second. Compare the magnitude of the force F with the weight of a pound.

63. Describe an experiment for demonstrating the principle of the parallelogram of forces to a class. A nail is driven into a wall and two strings are tied to its head. When the two strings are pulled horizontally and at right angles to one another with forces equal to 6 and 8 lbs. respectively, the nail can be dislodged. What forces would be needed if the strings were brought together and the nail pulled straight out? Illustrate your answer with a diagram.

64. Two forces, the magnitudes of which are proportional to the numbers 3 and 4, act on a point at right angles to each other. Draw a parallelogram as nearly to a scale as you can to show the direction and magnitude of the resultant, and deduce by measuring your diagram, or in any other way, the magnitude of the resultant.

65. What is meant by the parallelogram of forces? Give a diagram to illustrate your answer. Describe an experiment by means of which the truth of the proposition may be verified.

66. Two forces P and Q act upon a body. If P acted alone it would, in two seconds, produce in the body a velocity of 10 feet per second, while if Q acted alone it would, in three seconds, produce in the body a velocity of 18 feet per second. What velocities will P and Q produce in one second when acting together if the directions in which they tend to move the body are—(*a*) identical; (*b*) directly opposed?

Note.—For the sake of practice the student should work some of these examples, substituting centimetre-gram-second units for those given.

CHAPTER VI

PARALLEL FORCES AND CENTRE OF GRAVITY

Parallel Forces.—We have seen that the earth exerts a downward pull upon all objects on its surface, and that in consequence of this all things fall to the ground if unsupported. It follows, therefore, that everything which is supported above the earth's surface is constantly being pulled downwards, even though it does not fall. If a beam, for instance, is supported horizontally by resting the ends upon two posts, each particle of it may be regarded as being pulled earthwards by an attractive force. The direction of the pull is everywhere towards the centre of the earth, so for any one spot on the earth's surface we may consider the attractive forces due to gravity to be parallel to one another.

> Expt. 59.—Place the ends of a stiff lath or rod of uniform thickness upon two letter balances, or support the rod by hanging each end from a spring balance. Notice the weight borne by each balance; then weigh the rod, and so determine the proportion of the weight supported at each end.

This experiment represents on a small scale the case of a beam referred to before, and by using the spring balances we further see that the weight of a beam is equally divided between the two supports. In other words, we find that the two upward forces exerted by the balances are together equal to the downward force represented by the weight of the beam.

> Expt. 60.—Using the same arrangement as before, notice the reading shown by each balance when the lath is supported.

Then place a weight anywhere upon the lath and again observe the reading of each balance (Fig. 27). Repeat the experiment with the weight in different positions on the lath, and record the results in columns as shown below :—

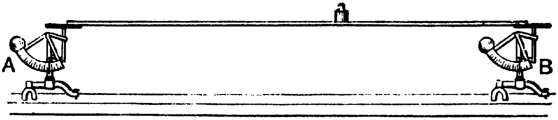

FIG. 27.—To illustrate Experiment 60.

Weight supported by A.	Weight supported by B.	Total readings A + B.	Weight of lath plus weight used.

If columns 3 and 4 are compared after performing an experiment of this kind they will be found to be practically the same, thus again showing that when the lath is in equilibrium the sum of the upward forces is equal to the sum of the downward forces. The following experiments further illustrate this principle.

EXPT. 61.—Suspend a light, stiff rod by a string which passes over a pulley and has attached to its other end a weight equal to that of the rod. The rod can then move as if it were weightless. Fasten a spring balance to a convenient point on the rod. Suspend weights A and B at the extremities of the rod so that they balance about this point, and show that the magnitude of the resultant is equal to the sum of the weights (Fig. 28).

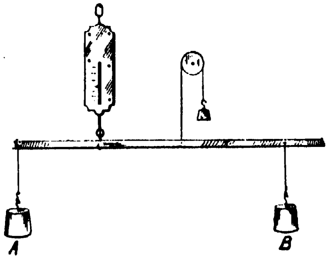

FIG. 28.—Parallel Forces.

Expt. 62.—Attach an additional weight to the string, which is fastened to the centre of the rod, and also weights to the ends, so that the rod is in equilibrium. Show that the pull on the balance is equal to the difference of the forces which act downwards and upwards respectively.

Principle of Parallel Forces.—We are now in a position to state definitely the principle of parallel forces demonstrated by the foregoing experiments. It may be expressed as follows:— **"The resultant of a number of parallel forces is numerically equal to the sum of those which act in one direction, less the sum of those which act in the opposite direction."** In other words, we may say, the resultant is equal to the *algebraic* sum of the forces

If two *equal parallel forces* act in the same direction upon a body, the total force will be obtained (as might be expected) by adding the two individual forces together. In like manner, if two *unequal parallel forces* act in opposite directions the net effect will be found by subtracting the lesser of the two forces from the greater, and the direction of the resultant will be that of the greater force.

Resultant of Parallel Forces.—It has been shown experimentally that the resultant of a system of parallel forces is equal in *magnitude* to the algebraic sum of the forces; the *direction* of the resultant is the same as that of the greater of the parallel forces. This is illustrated by Expts. 61 and 62; for while the forces due to the weights on the rod act vertically downwards, the resultant, represented by the pull of the spring balance, acts vertically upwards. The *position* which the resultant occupies with reference to the component forces can also be shown by the same experiments, or by the following one.

Expt. 63.—Using the arrangement described in Expt. 62, suspend different weights from the rod and then move the spring balance along the rod to a point about which the weights counterpoise one another. Repeat the experiment with the weights at unequal distances from the balancing point, and observe in each case (*a*) the reading of the spring balance, (*b*) the distance of each weight from the point at which the spring balance is attached to the rod. Record as below and compare columns 5 and 6.

Wt. A.	Wt. B.	Dist. Ac.	Dist. Bc.	A×Ac.	B×Bc.

The first part of a record of this kind will show that the magnitude of the resultant force is equal to the sum of the components, while the second part will prove that in each case when the rod is in equilibrium one weight multiplied by its distance from the point of action of the resultant is equal to the other weight multiplied by the other's distance from the resultant. Or, expressing the result as an equation, and substituting the word force for weight, we have :—

$$\frac{\text{Force on}}{\text{one side}} \times \frac{\text{Distance from}}{\text{Resultant}} = \frac{\text{Force on}}{\text{other side}} \times \frac{\text{Distance from}}{\text{Resultant.}}$$

If the component forces are equal, their distances from the resultant will also be equal; and if they are unequal the resultant will always be nearer to the greater force. In other words, a small force is at a large distance from the resultant, and a large force is at a small distance.

This follows from what has been already said (p. 49) concerning the moments of forces tending to turn a body in opposite directions. It will be remembered that the moment of a force is obtained by multiplying the force (measured in units) by the perpendicular distance between the point at which it acts, and the fulcrum or hinge about which the body turns. If a rigid object such as a lever or balance, capable of turning about a fixed point in one plane is at rest, then the sum of the moments of the forces tending to turn it in one direction is equal to the sum of those tending to turn it in the opposite direction. Expressed as an equation, we have :—

$$\frac{\text{Sum of moments}}{\text{on one side of fulcrum}} = \frac{\text{Sum of moments}}{\text{on other side of fulcrum.}}$$

Still another way of regarding this important truth is to say that the *algebraic sum of the moments is always equal to* 0.

Conditions for the Equilibrium of Three Parallel Forces.—The student should now be in a position to understand clearly the conditions which must prevail for three parallel forces to be in equilibrium.

Expt. 64.—Suspend a light rod horizontally from two spring balances (Fig. 29), and hang a weight from any point between the balances.

Observe (*a*) the readings of the balances.

(*b*) The length of the rod between the balances.

(*c*) The length of rod between the weight and the balance on each side.

Vary the distance of the weight, and again make observations. Also repeat the experiment with a different weight.

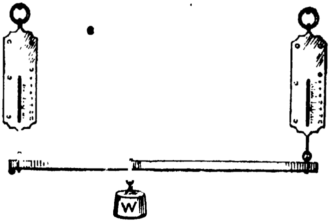

Fig. 29.—Equilibrium of Three Parallel Forces.

Neglecting the weight of the rod, this experiment is a demonstration of the conditions for the equilibrium of three parallel forces. The weight will be found to be equal to the sum of the readings of the spring balances. In other words, the resultant force will be found equal to the sum of the components. Moreover, the readings of the spring balances will be found to be in inverse ratio to the distances of the balances from the weight; while the weight itself will be proportional to the length of the rod between the balances. The relation is shown graphically by Fig. 30, in which the three forces are represented by the lines, CW, AD, BE. The forces bear exactly the same proportion to one another as the distances AB, BC, CA. We have, therefore, as the conditions of equilibrium of three parallel forces the following expression:—

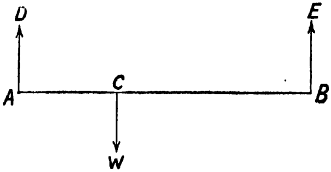

Fig. 30.—Equilibrium of Parallel Forces.

$$\frac{\text{Force CW}}{\text{Distance AB}} = \frac{\text{Force AD}}{\text{Distance BC}} = \frac{\text{Force BE}}{\text{Distance CA}}$$

Or, expressed in words, *each force is proportional to the distance between the other two forces.*

Example.—A weight of 120 lbs. is suspended from a light rod (whose weight may be neglected), the ends of which rest upon the shoulders of two men, A and B (Fig. 31). The rod is 6 feet long, and the weight is 2 feet from the man A, and 4 feet from the man B. What proportion of the weight is borne by each man?

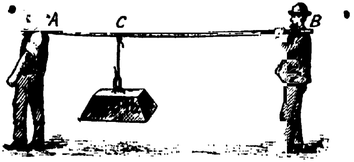

Fig. 31.—Relation of Parallel Forces to one another.

From the above we see that

$$\text{Amount borne by A} : \text{Amount borne by B} = \text{B's distance from weight} : \text{A's distance from weight}$$
$$= 4 : 2 = 2.$$

So that the man A bears twice as much as the man B, that is, he bears 80 lbs. and B carries 40 lbs.

The Principle of Moments applied to Parallel Forces.—It will be evident that the preceding problem resolves itself into one on the moments of forces. Suppose we consider the shoulder of the man A to be a fulcrum about which the rod may turn, then, since the rod is in equilibrium, the moments of the forces tending to turn it are equal and opposite. We may, therefore, write :—

$$\text{B's exertion} \times \text{distance from A} = \text{weight} \times \text{distance from A.}$$
$$\text{Or, B's exertion} \times \text{AB} = 120 \times 2.$$
$$\text{That is, B's exertion} \times 6 = 240$$
$$\text{Therefore, B's exertion} = \frac{240}{6} = 40 \text{ lbs}$$

The force exerted by A may be found in a similar way by taking the moments about the shoulder of B considered as a fulcrum. In this case we have

$$\begin{aligned} \text{A's exertion} \times \text{distance from B} &= \text{weight} \times \text{distance from B.} \\ \text{Or, A's exertion} \times 6 &= 120 \times 4. \\ \text{Therefore, A's exertion} &= \frac{120 \times 4}{6} = 80 \text{ lbs.} \end{aligned}$$

We may also consider the moments of forces acting at the point C, from which the weight is suspended. Then we have

$$\begin{aligned} \text{A's exertion} \times AC &= \text{B's exertion} \times BC \\ \text{Or, A's exertion} \times 2 &= \text{B's exertion} \times 4 \\ \text{Therefore, A's exertion} &= \frac{\text{B's exertion} \times 4}{2} \\ &= \text{B's exertion} \times 2 \end{aligned}$$

If B's exertion is denoted by x, then A's exertion is expressed by $2x$. The total weight supported is 120 lbs.; so that

$$\begin{aligned} \text{A's exertion} + \text{B's exertion} &= 120 \text{ lbs.} \\ \text{Or } 2x + x &= 120 \text{ ,,} \\ \text{That is } 3x &= 120 \text{ ,,} \\ \text{Therefore } x &= \frac{120}{3} = 40 \text{ lbs.} \\ &= \text{B's exertion.} \end{aligned}$$

We have thus taken moments about three different points and have found that they all give the same results. As a matter of fact, it does not matter what point is considered as the fulcrum when forces act upon a body in one plane, in the manner here considered. Under these conditions the rule already given for the moments of forces holds good, and the moments may be taken about any point in the plane in which they act. As it is important that the student should thoroughly understand this, we give another example :—

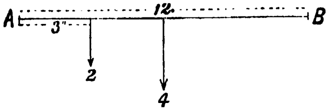

FIG. 32.—To illustrate the Principle of Moments applied to Parallel Forces.

$$\frac{\text{Force CW}}{\text{Distance AB}} = \frac{\text{Force AD}}{\text{Distance BC}} = \frac{\text{Force BE}}{\text{Distance CA}}$$

Or, expressed in words, *each force is proportional to the distance between the other two forces.*

EXAMPLE.—A weight of 120 lbs. is suspended from a light rod (whose weight may be neglected), the ends of which rest upon the shoulders of two men, A and B (Fig. 31). The rod is 6 feet long, and the weight is 2 feet from the man A, and 4 feet from the man B. What proportion of the weight is borne by each man?

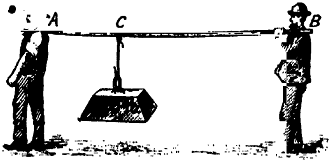

FIG. 31.—Relation of Parallel Forces to one another.

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$$\text{That is, B's exertion} \times 6 = 240$$
$$\text{Therefore, B's exertion} = \frac{240}{6} = 40 \text{ lbs}$$

forces is represented by the line GF, and the centre of the forces is the point G. The point G, through which the resultant (GF) of the parallel forces due to the weights of the individual particles of the stone passes, is known as the **centre of gravity**. For the stone to be in equilibrium, the string must be attached to a point in the line GF, or GF produced.

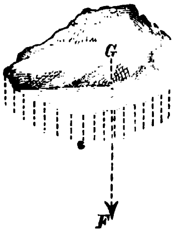

FIG. 33.—Parallel Forces due to Gravity.

Every material object has a centre of gravity, and the position of this point for a particular object is the same so long as the object retains the same form.

Experimental Methods of Determining Centre of Gravity.—The centre of gravity of such geometrical figures as circles, squares, and parallelograms is really the centre of the figures, and can therefore be determined geometrically. In the case of unsymmetrical figures, however, the centre of gravity cannot be so easily found by geometry, and is best determined by experiment. The following experiments illustrate the method employed, and the principle involved :—

EXPT. 65.—Procure an uncut pencil. Tie a piece of thread round the pencil and adjust the thread in such a position that the pencil is suspended horizontally from it. Measure the distance of the thread from each extremity of the pencil.

The pencil may be regarded as a straight line, and the observations will show that the centre of gravity is in the middle of it.

EXPT. 66.—Obtain several sheets of wood, zinc or cardboard cut into various shapes as shown in Fig. 34, and drill holes in the positions indicated. Using one at a time, tie pieces of string to the sheet, passing each through one of the holes. Hang the plate by one of the strings to a support such as one of the rings of a retort stand. Allow it to come to rest and draw a chalk mark across the plate in the same straight line with the string as shown by the dotted line in the figure. Now attach the same plate by one of the other threads exactly as before and again make a mark in continuation of the string. The two chalk marks intersect at a point marked G, Fig 35. Untie and do the same with another string, the third line passes through the intersection of the first two. Obtain a similar point for each of the other plates. Also determine in the same way a similar point for irregular plates of wood, zinc or cardboard.

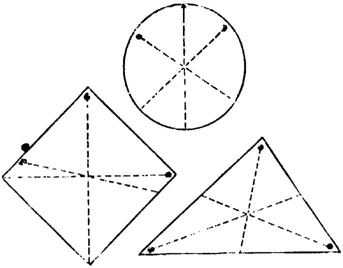

FIG. 34.—Figures for the Determination of Centres of Gravity.

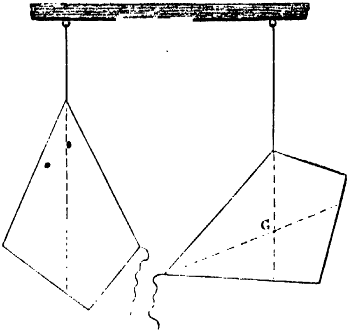

FIG. 35.—The Determination of Centre of Gravity.

EXPT. 67.—Balance the pieces of wood, zinc, or cardboard, one after the other, upon a pointed upright, at the intersection of the chalk marks; each will be found to set itself in a horizontal position. If the plates are of cardboard, a better plan is to make a pin-hole at the centre of gravity and pass a thread knotted at one end through it. The plate can then be held up and will be found to set horizontally.

EXPT. 68.– Determine experimentally the centre of gravity of a plate in the form of a parallelogram. Mark the point with a pencil; then turn over the plate and draw the two diagonals upon the opposite side. Make a pin-hole where the diagonals intersect. The point where the diagonals intersect will be found to be practically the same as the centre of gravity.

Centre of Gravity of a Triangular Plate.—EXPT. 69.—Repeat the preceding experiment, with a triangle cut out of cardboard. After finding the centre of gravity, prick a pin-hole through the cardboard, then turn over the triangle and draw a line from each angle through the pin-hole to the opposite edge. Now, taking each edge in turn as the base of the triangle, determine (*a*) the lengths of the two parts into which each base is divided by the lines drawn, (*b*) the proportion which the distance of the hole from each base bears to the length of the line from that base to the opposite angle.

It will be found as the result of this experiment that the line from each angle to the base divides the base into two equal parts, and also that the distance of the hole from the base is one-third the whole length of the line. We may, in fact, consider a triangular plate as made up of a number of narrow strips of material which decreases in length from the base to the apex. The centre of gravity of each strip is the middle of the strip; hence the line drawn from the apex to the middle of the base passes through each centre of gravity. By taking another side

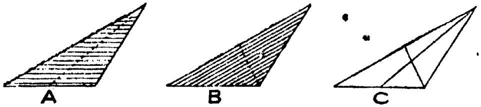

FIG. 36.—Geometrical illustration of Centre of Gravity of a Triangular Plate.

as base a similar line can be drawn from the middle to the opposite angle. These lines intersect at one-third the distance up, measured from the base, and the point of intersection is the centre of gravity of the triangular plate (Fig. 36).

Centre of Gravity of a Quadrilateral.—EXPT. 70.—Cut a four-sided figure out of cardboard, and draw a line connecting two of the opposite angles. Find by bisecting this line and taking one-third the distance from the middle to the opposite angle, the centre of gravity of each of the triangles into which the figure is divided. Connect the two points found. Then draw the other diagonal of the quadrilateral, repeat the measures, and connect the centres of gravity as before. Make a hole where this short line cuts the other, and show, by passing a piece of knotted thread through it, and so suspending the cardboard, that the point determined in this way is the centre of gravity of the whole figure.

Other Centres of Gravity.—EXPT. 71.—Procure a skeleton cube or tetrahedron, and suspend it as in the preceding experiments. Mark the verticals through the point of suspension by light wires attached by wax, and thus find the position of the centre of gravity.

The centre of gravity of a skeleton cube may be found in this way to be the intersection of the diagonals. In a similar manner, the centre of gravity of a right cylinder may be shown to be the middle point of the axis.

EXPT. 72.—Find the centre of gravity of an open wicker-work basket, such as a waste-paper basket. To do this, suspend the basket, and hang a plumb-line from the point of suspension. Tie a piece of thread across the basket in the direction of the plumb-line; then suspend the basket from another point, and notice where the plumb-line crosses the thread. The point of intersection is the centre of gravity.

Equilibrium.—When a body is at rest all the forces acting upon it balance one another (or, what is the same thing, any force is equal and opposite to the resultant of the remaining forces) and it is said to be in equilibrium. It is in *stable equilibrium* when any turning motion to which it is subjected raises the centre of gravity; in *unstable equilibrium* when a similar movement lowers the centre of gravity, and in *neutral* equilibrium when the height of the centre of gravity is unaffected by such movement. Consequently, if a body in stable equilibrium is disturbed, it returns to its original position; if in unstable equilibrium, it will, if disturbed, fall away from its original position; while if the condition of equilibrium is neutral it will, under similar circumstances stay where it is moved to.

EXPT. 73.—Procure an oblong strip of board or cardboard (Fig. 37). Bore a hole through the oblong near one end, and one through the middle. Support the strip as at A; it is then in stable equilibrium, for the slightest turn either to right or left raises the centre of gravity. When supported as at B, the strip is in neutral equilibrium; and when supported as at C it is in unstable equilibrium.

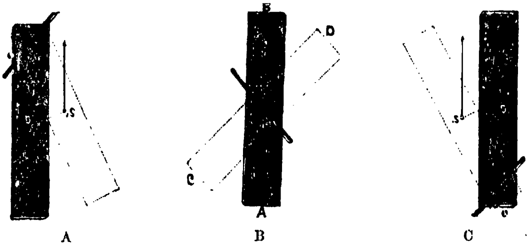

FIG. 37.—Conditions of Equilibrium.

A ball resting upon a table is in neutral equilibrium; for when it rolls the height of the centre of gravity is not changed.

Conditions of Stability.—EXPT. 74.—See whether any of the cardboard figures used in the preceding experiments can be hung loosely from a pin with the centre of gravity above the pin. Try the experiment with the pin at different distances from the centre of gravity. In each case notice the direction of the line connecting the centre of gravity with the point of support when the figure comes to rest.

This experiment illustrates the conditions with reference to centre of gravity and point of support which must be fulfilled in order that an object shall be in equilibrium. It will be found that the centre of gravity is in every case below the point of support when the suspended object is in equilibrium. The greater the distance between the point of support and the centre of gravity, the greater is the tendency to return to the position of equilibrium.

When the centre of gravity and the point of support of an object are close together the equilibrium of the object is easily disturbed. A good balance partly owes its sensitiveness to this condition, the centre of gravity and point of support being designedly brought close together.

It has been shown that in the case of a freely suspended object the centre of gravity is at its lowest point when the object is in equilibrium. Let us see how this applies to a body supported upon a surface below the centre of gravity.

Equilibrium of Object resting upon a Base.—EXPT. 75.—Procure an oblong block of wood of about the same size as this book, and about an inch thick. Draw the two diagonals on one

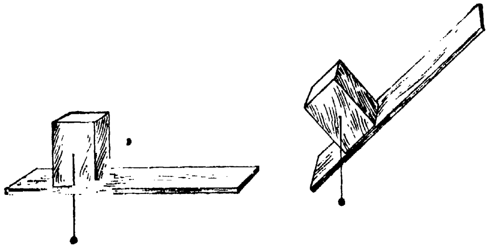

FIG. 38.—Relation between Centre of Gravity and the Base of Support.

of the faces; the centre of gravity of the block will be inside the block at a point below the intersection of the diagonals. Press a pin partly into the block at the point where the diagonals intersect, and suspend from it a small plumb-line made of a short piece of thread and a bit of lead (Fig. 38). Now place the block upon a board, and notice the direction of the plumb-line. Tilt the board until the block topples over. When this happens, notice the direction of the plumb-line with reference to the base.

It will be found by this experiment that the block topples over as soon as the plumb-line falls beyond the point A of the base. For any object resting upon a base to be in equilibrium, a vertical line drawn from the centre of gravity downward must fall within the base. When this vertical line falls outside the base, the body topples over.

EXPT. 76.—Repeat the preceding experiment with thick blocks of wood triangular in shape, and with cylinders and cones. In this way show that a body is only in equilibrium when the centre of gravity is vertically above some point of the supporting surface.

It will be evident from these experiments that a body is least liable to be upset when the centre of gravity is at a considerable distance from all parts of the edge of the base; for when this is the case the body has to be tilted through a large arc before the centre of gravity falls outside the base.

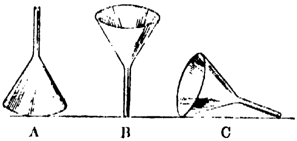

FIG. 39.—A funnel in (A) Stable Equilibrium, (B) Unstable Equilibrium, (C) Neutral Equilibrium.

A funnel standing upon its mouth is an example of a body which cannot be easily overturned on account of the low centre of gravity and its distance from the edge of the base (Fig. 39). It is then in stable equilibrium. If the funnel is stood upon the end of the neck it can easily be overturned, because very little movement is required to bring the centre of gravity outside the base. It is then in unstable equilibrium. When the funnel lies upon the table it is in neutral equilibrium, for its centre of gravity cannot then get outside the points of support.

CHIEF POINTS OF CHAPTER VI.

Parallel Forces.—The resultant of a number of parallel forces is numerically equal to the sum of those which act in one direction, less the sum of those which act in the opposite direction. In other words the resultant of a system of parallel forces is equal in *magnitude* to the algebraic sum of the forces; the *direction* of the resultant is the same as that of the greater of the parallel forces.

The Centre of Gravity of a rigid body is the point upon which the body could be supported or balanced; in other words, it is the point through which the resultant of the parallel forces due to the weights of the individual particles passes.

Every material object has a centre of gravity, and the position of this point for a particular object is the same so long as the object retains the same form.

The centre of gravity of such geometrical figures as circles, squares and parallelograms is the centre of the figures, and can be determined geometrically. The centre of gravity of unsymmetrical figures can be determined by experiment.

The centre of gravity of a triangle is located on the line drawn from one of the angles to the middle point of the side opposite and at a distance of one-third of this line's length from that side of the triangle.

Equilibrium.—A body is said to be in equilibrium when all the forces acting upon it balance one another. It is in **stable equilibrium** when any turning motion to which it is subjected raises the centre of gravity; in **unstable equilibrium** when a similar motion lowers the centre of gravity; and in **neutral** equilibrium when the height of the centre of gravity is unaffected by such movement.

QUESTIONS ON CHAPTER VI.

1. Describe an experiment to prove that the resultant of a number of parallel forces is numerically equal to the sum of those which act in one direction, less the sum of those which act in the opposite direction.

2. State the conditions for the equilibrium of three parallel forces. Describe an experiment which shows these conditions.

3. Apply the principle of moments to explain the conditions of equilibrium for parallel forces. Give an example.

4. What do you understand by the centre of gravity of a body?

5. Describe an experimental method for finding the centre of gravity of any body, *e.g.*, a waste-paper basket.

6. Where is the centre of gravity of a triangle located.

7. How could you find by a geometrical construction the centre of gravity of a circular plate, a square piece of cardboard, or any other geometrical figure?

8. When is a body said to be in equilibrium? Distinguish between

stable, unstable, and neutral equilibrium. What is the condition which determines the nature of the equilibrium?

9. Describe an experiment which exemplifies each of the terms in the preceding question.

10. Give instances of bodies in stable, unstable, and neutral equilibrium.

11. How would you determine the centre of gravity of an iron hoop made by joining two semi-circles, one thicker than the other? Explain how the observations could be used to find out which was the thicker half of the hoop.

CHAPTER VII

PRINCIPLE OF ARCHIMEDES, AND RELATIVE DENSITY

Principle of Archimedes.—Every one has noticed, when in water, that the body is buoyed up and appears lighter than when on land. The liquid appears to exert an upward force of buoyancy which partly counteracts the weight of the body. The force which a liquid exerts upon an object immersed in it is measured in the following experiments:—

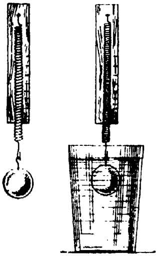

FIG. 40.—Loss of Weight in Water.

EXPT. 77.—Suspend from a spring balance a metal ball or cube or cylinder, the volume of which is known or can be calculated from the dimensions. Or, instead, use any object, such as a glass stopper, the volume of which has previously been determined by displacement of water (see Expt. 43). Notice the weight of the object in air, when suspended from the balance; then immerse the ball in a glass of water and observe the apparent loss of weight (Fig. 40).

Weight in air (grams).	Weight in water (grams).	Difference (grams).	Volume in cubic centimetres.

The number of cubic centimetres in the object is also the number of cubic centimetres of water displaced by the object; and, since one cubic centimetre of water weighs one gram (the weight varies slightly at different temperatures, but we may disregard the variation here), the number of grams which the water displaced weighs is the same as the number of cubic centimetres displaced. The experiment will show that the loss of weight of an object submerged in water is equal to the weight of the water displaced by the object; or, expressed in another way, *the upthrust experienced by an object immersed in water is equal to the weight of the water displaced.*

EXPT. 78.—Repeat Expt. 77 with a pair of scales instead of a spring balance. To do this, suspend the object by means of a piece of thread from a hook above the left pan of a balance, and weigh it. Then place a glass of water upon a platform or stage so that the object is immersed as in Fig. 41, and weigh again. The loss of weight will, as before, be found to be equal to the weight of the volume of water displaced by the object used.

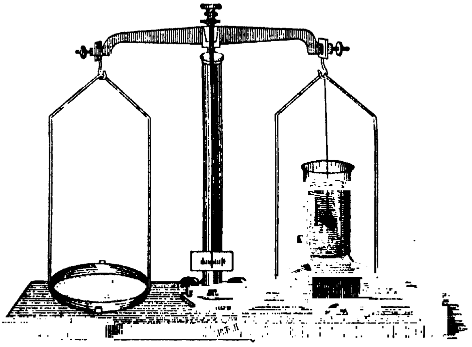

FIG. 41.—Determination of the Loss of Weight of a Body in Water. (*But notice the stopper is suspended from the wrong arm, see p.* 51.)

EXPT. 79.—Using the graduated glass jar already described (Expt. 43), fill it about two-thirds full of water, noticing the mark level with the top of the water. Procure a cylindrical tin canister about half the diameter of the jar. Place the canister in the water, and gradually put shot into it until it just sinks in the water when the cover is on. Pour the water displaced by the canister into a beaker counterpoised upon a balance; then take out the canister, wipe it, and place it in the other pan of the balance. You will find that the weight of the canister is practically the same as the weight of the water displaced.

EXPT. 80.—Repeat the preceding experiment, using another liquid, such as methylated spirit or turpentine instead of water.

The foregoing experiments justify the following conclusion—known as the "Principle of Archimedes": **When a body is submerged in a liquid it loses weight equal to the weight of the liquid displaced by it.** This principle applies to all bodies, whether they are lighter or heavier than the liquid in which they are immersed.

So far we have only referred to bodies which sink, or remain suspended, in the liquids displaced, and it is now necessary to deal with bodies which float.

Experiments on Flotation.—EXPT. 81.—Procure a rectangular rod of wood or a pencil, and weigh it. Select a narrow glass jar graduated into cubic centimetres, or a burette; fill it with water up to a certain mark, and put the rod into it.

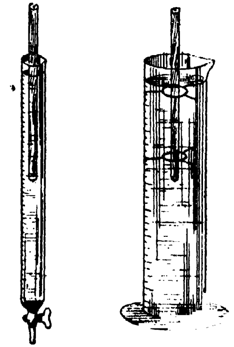

FIG. 42.—Flotation of Wooden Rod in a Burette and in a Graduated Jar.

(Fig. 42). Notice the length of rod immersed and how many cubic centimetres or grams of water are displaced by the rod. Repeat the experiment with turpentine, brine, and methylated spirit, in each case running off the liquid displaced and weighing it. Record your observations as below :—

Length of immersed portion of rod... ... ... cm.
Length of whole rod ... ... cm.
Weight of rod ... ... gm.
Weight of liquid displaced ... ... gm.

A comparison of the results thus obtained will show that the weight of the whole rod is equal to the weight of the liquid displaced by the portion of the rod immersed.

Expt. 82.—Procure a wooden cube or rectangular block. Slightly oil the block and place it in a glass of water. Mark upon it with a pencil where the surface of the water touches it. Then take out the block and determine the volume of the immersed portion in cubic centimetres. You will then know the volume, and therefore the weight, of the water displaced. Record as below :—

Weight of block ... ... gm.
Volume of immersed portion of block ... c.c.
Therefore weight of water equal to
the immersed portion of block ... ... gm.

In this case, again, it will be found that the weight of the whole block is equal to the weight of the volume of water displaced by the portion immersed.

Expt. 83.—Procure a large wooden pill-box, and float it, by adding shot, without the cover, in water, up to a mark which you have made beforehand upon the curved surface. Show, as before, that the weight of the volume of water displaced is equal to the weight of the whole box and the shot. Repeat the experiment by putting more shot in the box, and so making it sink deeper in the water.

Expt. 84.—Fill the graduated jar with water up to a certain mark, and then float a test-tube in it. Put sufficient shot into the test-tube to make it float upright ; and, when it does so, notice the volume of water above the initial level of the water in the jar. The weight of this water in grams will be equal to the number of cubic centimetres above the initial level. Take out the test-tube and weigh it; the weight will be found equal to the weight of water displaced. Repeat the experiment with the test-tube floating at different depths.

Applications of Archimedes's Principle.—From the foregoing experiments it is easy to understand many interesting facts. For instance, a ship made of iron and containing all kinds of heavy things is able to float in water although the material of which they are made is heavier than water. Evidently the reason is that the ship and all its contents only weighs the same as the volume of water displaced by the immersed part of the hull. Or, to put the fact another way, the ship as a whole weighs less than a mass of water the same size as the ship would weigh. We have seen (Expt. 79) that when an object weighs the same as an equal volume of water, it will remain suspended in the water; when it weighs more than an equal volume of water it sinks (Expt. 78); and when it weighs less than an equal volume of water it floats (Expts. 81 and 82). This principle applies to all fluids, that is, all liquids, and to all gases as well. It explains that a balloon rises because the gas contained in it, together with the bag and all the tackle, weighs less than the weight of an equal volume of air. If the balloon were free to ascend it would rise to a height where its weight would be equal to the weight of an equal volume of the surrounding air.

Density.—We shall now apply the principle demonstrated by the foregoing experiments to the determination of the densities of solids and liquids.

EXPT. 85. Procure equal volumes of different substances, *e.g.*, a cubic inch of wood, lead, cork, marble, and determine their masses by means of a balance. Notice they are different.

EXPT. 86. Compare the volume of a pint of water with that of one and a quarter pounds of iron. Observe their masses, as determined by a balance, are equal, but their volumes very unequal.

EXPT. 87. Fill two equal flasks with water and methylated spirit respectively, and weigh. Show that the masses are different.

We thus see that equal volumes of different substances have different masses. This truth is expressed by saying they have different densities.

If we keep to the unit of volume, the numbers representing the masses of this volume of the kinds of matter we experiment upon are a direct measure of the densities of these bodies. We can thus define it :—**Density is the mass of a unit volume of a substance.** It follows from this definition that if the volume of a body is multiplied by its density, we shall obtain its mass

$$\textbf{volume} \times \textbf{density} = \textbf{mass}$$

$$\text{or density} = \frac{\text{mass}}{\text{volume}}.$$

In using this relation between the volume and mass care must be taken to use the proper units. In all scientific work it is customary to adopt the cubic centimetre and gram as the units of volume and mass respectively.

Density can be regarded in another way. It is clear that if we pack twice the amount of mass into a given volume we shall have doubled its density, so that density may be looked upon as the closeness with which a mass is packed into a given volume.

When we use the cubic centimetre and gram as the units of volume and mass and apply the equation given above, the density of water at 4° works out to be 1 ; for all other forms of matter the number will be either a fraction or multiple of this value.

Relative Density.—Hence the ratio of the weight of *any volume* of a substance to the weight of the same volume of water at 4° C. is equal to the absolute density of the substance, while, since the density of water only changes slightly with the temperature, the ratio of the weight of the substance to the weight of water at temperatures other than 4° is a number very nearly equal to its absolute density. This ratio is the **relative density** of the body, or, as it is frequently called, the **specific gravity.**[1]

When relative densities are used, care must be taken to indicate the temperature of the water to which its weight is referred. Thus, *relative density alcohol* $^{15°}/_{15°}$ indicates the ratio of the weight of equal volumes of alcohol and water, both at 15°; while if we wrote $^{15°}/_{4°}$ it would indicate the ratio of the weight of any volume of alcohol at 15° to that of the same volume of water at 4°.

The student should carefully notice the difference between the meanings of the terms *density* or *absolute density* and *relative density*: the first is dependent on the units of mass and volume, the second being merely a numerical ratio does not vary with these units.

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[1] The term specific gravity is so generally used in this sense that this meaning is here given. It is, however, preferable to restrict the term to its true meaning, *i.e.*, the *weight* of unit volume, and to always use the term *relative density* when the above ratio is meant.

Hydrostatic Method of Determining Relative Densities.—It has already been pointed out, and illustrated by experiment, that equal volumes of different materials have different masses, which truth is expressed by saying that they have different absolute densities. Again, the experiments on the principle of Archimedes have shown that some substances are heavier and some lighter than an equal volume of water. Let us use the principle to find exactly how much heavier or lighter certain substances are than an equal volume of water, that is, their *relative densities* compared with water as a standard.

There are two numbers which we wish to determine:—(i) The weight of the body of which the density is required, *i.e.*, its weight in air; (ii) the weight of an equal volume of water.

The first number can be obtained directly by hanging the body from a hook at the top of one scale pan, and then placing accurate weights on the other, until the weight of the body is exactly balanced. This only requires practice to be able to accomplish it with the greatest precision. To find the second number, a glass of water is placed upon a small platform (Fig. 41) so that the body under experiment is immersed in it. The loss of weight then experienced is equal to the weight of an equal volume of water. This experiment shows, therefore, the weight of the body and the weight of an equal volume of water. The proportion between these two numbers, that is, the first divided by the second, is the relative density of the body. For,

$$\text{Relative density} = \frac{\text{weight of substance}}{\text{weight of equal vol. of water}}.$$

And as

$$\text{loss of weight in water} = \text{weight of equal vol. of water}$$

we can put down

$$\textbf{Relative density} = \frac{\textbf{weight of substance}}{\textbf{loss of weight in water}}$$

Expt. 88.—Find by weighing, and determining the loss of weight in water, the relative densities of a halfpenny, a shilling, and a sovereign.

Expt. 89.—Find, by the method of the preceding experiment, the relative densities of two or three common solids, such as brass, sulphur, copper, zinc, lead, glass, coal, and flint.

EXPT. 90.—(1) Weigh a piece of india-rubber. Cut a strip of lead, or of other metal, sufficient to act as a sinker to the india-rubber, and weigh it (2) in air and (3) in water. Attach the india-rubber to the lead, and find (4) the combined weight in water.

Then from 2 - 3 you know the weight of the volume of water equal to the volume of lead, and from (1+2) - 4 the weight of the volume of water equal to the combined volume of the lead and rubber, and hence the difference between these values gives the weight of the volume of water equal to the volume of the rubber.

EXPT. 91.—Repeat the preceding experiment with different kinds of wood, cork, and other solids which float in water, using a piece of lead as a sinker in each case.

EXPT. 92.—Weigh a glass stopper in air, then immerse it successively in water, turpentine, methylated spirit, olive oil, and petroleum, and notice the loss of weight in each case. The loss of weight experienced by the glass stopper in each experiment is equal to the weight of a portion of liquid of the same volume as the stopper. The numbers obtained therefore represent the weights of equal volumes of water, turpentine, methylated spirit, olive oil, and petroleum, and by dividing each by the number obtained in the case of water, the relative densities of the liquids are obtained.

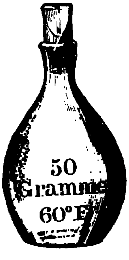

FIG. 43.—A Relative Density or Specific Gravity Bottle.

Determination of Relative Densities by Relative Density Bottle.—The method of determining the relative densities of liquids by means of the principle of Archimedes is an indirect one; a simpler plan is to use a specific gravity or relative density bottle. Such a bottle (Fig. 43) consists of a small glass flask, holding about 50 grams of water. It is provided with a nicely-fitting ground stopper, which is in the form of a tube with a very small bore through it. It is used in determining the relative density of liquids and powders. To use it, we must first know the weight of the empty bottle and stopper. The bottle is then filled with pure water, the stopper inserted, and the water which is forced through the hole in the stopper wiped off, and the bottle and its contents weighed. In this way the weight of water which just fills the bottle is found. If now we empty the bottle and carefully dry it inside and out, and fill it with the liquid of which the density is required, say spirits of wine, and weigh again, we have the weight of the liquid which just fills the bottle, or the weights of equal volumes of the liquid and water, the proportion of which gives us the relative density of the liquid. Instead of a bottle of this kind, a flask having a narrow neck around which a mark has been made may be used, or, any bottle with a vertical file mark made on the stopper, will do. The weight of water which fills the bottle up to the mark may thus be compared with the weight of liquid which fills it to the same mark.

Expt. 93.—Counterpoise an empty specific gravity bottle, or a flask having a mark on its neck. Fill the flask up to the mark with methylated spirit and weigh it; then empty the flask, dry it, and fill with water up to the same mark. Weigh again, and from the two weights find the relative density of the spirit, remembering that

$$\text{Relative density} = \frac{\text{weight of substance}}{\text{weight of equal vol. of water}}$$

Expt. 94.—Following the method of the previous experiment, determine the relative densities of two or three liquids, such as turpentine, milk, vinegar, beer, wine, sea-water or a solution of salt, and ink.

Expt. 95.—Weigh out about 100 grams of shot. Fill the specific gravity flask with water, and counterpoise it together with the shot. Next put the shot into the bottle, and remove the water displaced. Add weights until the index of the balance swings evenly. The weights added must equal the weight of the water displaced, that is, the weight of a volume of water equal in volume to the shot. Therefore—

$$\text{Relative density of the shot} = \frac{\text{weight of shot}}{\text{weight of water displaced}}.$$

Expt. 96.—Find by the method used in the preceding experiment the relative densities of such common things as tin tacks (which are really made of iron), bits of slate pencil, brass wire, and brass nails.

Determination of Relative Densities by a U-tube.—A graphic method of showing the relative densities of liquids is obtained by means of a glass tube bent in the form of a **U**, and therefore called a **U**-tube.

Expt. 97.—Cut off two pieces of glass tube, each about 30 cm. long; connect the tubes with india-rubber tubing about 18 cm. long, and fix them upright upon a strip of wood. Pour mercury into one of the tubes until it reaches a horizontal line drawn a little above the junctions (Fig. 44). Now introduce water into one of the tubes by means of a pipette, and notice that the mercury on which the water rests is pushed down; afterwards introduce sufficient water into the other tube to bring the mercury back to its original level. The length of each column of water will be found the same. Repeat the experiment with varying amounts of water.

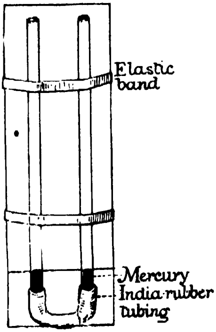

Fig. 44.—Determination of Relative Densities by a **U**-tube.

The mercury in the bend of the **U**-tube evidently acts as a balance, which enables us to balance columns of different liquids in the upright arms.

Expt. 98.—Nearly fill one of the tubes with methylated spirit, and balance it with water introduced into the other tube. Measure the lengths of the two columns.

As these two lengths of liquid balance one another it will be evident that the shorter of the two columns, namely, the water column, has a greater relative density than the longer column.

If a column of liquid 40 cm. long balance a similar column of water of half that length, would the liquid be double or half as dense as the water?

By thinking over this question and the result of the experiments you will be able to understand that the densities of two liquids balanced in a **U**-tube are in inverse proportion to the lengths of the columns; in other words, the density of one liquid is less than the density of the other in proportion as its column is longer than the balancing column of the other. If water is used in one of the tubes, you can determine the relative density of the other liquid as below.

Density of spirit : density of water : : water column : spirit column.

$$\text{Therefore relative density of spirit} = \frac{\text{water column}}{\text{spirit column}}.$$

Hare's Apparatus for Determination of Relative Densities.—A convenient arrangement to use instead of a U-tube to determine the relative densities is represented in Fig. 45, and is known as Hare's apparatus. Two straight glass tubes are connected at the top by a three-way junction, upon the unconnected end of which a piece of india-rubber tubing is placed. The lower ends of the tubes dip into beakers containing the liquids the relative densities of which have to be determined. By applying suction to the free end of the rubber tube, the two liquids are drawn up the glass tubes, and the heights of the liquid columns above the level of liquids in the beakers will be inversely proportional to the relative densities of the liquids employed. The principle is thus precisely the same as that of the U-tube, but by using the form of apparatus here described, the relative densities of liquids which mix can be more conveniently found than by the ordinary U-tube in which the liquids have to be poured.

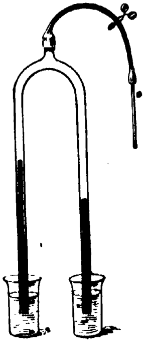

FIG. 45. Hare's Method of Determining Density.

Results of Accurate Determinations of Relative Densities.—When exact determinations of relative densities are made, it is necessary to take into consideration the fact that the density of water varies with the temperature, for reasons which will be understood later. The following table shows the relative densities of a number of solids and liquids :—

TABLE OF RELATIVE DENSITIES.

Solids and Liquids at 0° C. compared with water at 4° C.

Solids.

Platinum (rolled)	22·069
Gold (stamped)	19·362
Lead (cast)	11·352
Silver (do.)	10·474
Copper (do.)	8·788
Steel (not hammered	7·816
Iron (cast)	7·207
Heavy spar (barium sulphate)	4·430
Diamond	3·531 (varies)
Marble	2·837
Rock crystal	2·653
Anthracite	1·800
Coal (compact)	1·329
Melting ice	0·930
Oak	0·845
Cork	0·240

Liquids.

Mercury	13·598
Bromine	2·960
Sea-water	1·026
Distilled water (4° C.)	1·000
Olive oil	0·915
Alcohol (absolute)	0·793

Relative Density of Gases.—The density of a gas is not compared with that of water as a standard, because the number which would be obtained by such a comparison would be so exceedingly small. The standard density adopted is that of the gas hydrogen, which is the lightest form of matter known. The densities of other gases will in every case therefore be greater than unity. To find the density of a gas, all that it is necessary to do is to weigh a flask filled with pure dry hydrogen under certain conditions of temperature and pressure (*see* Chap. VIII), and then to weigh the same flask filled with the gas of which the density is required under the same conditions of temperature and pressure; the number of times the gas is heavier than hydrogen is the number representing its density. There are several important precautions which have to be adopted, but for an account of these we must refer the student to works on chemistry.

RELATIVE DENSITIES OF GASES.

Hydrogen	1·0
Nitrogen	13·9
Oxygen	15·9
Chlorine	35·2
Ammonia	8·45
Steam	8·95
Carbon monoxide	13·90
Carbon dioxide	21·85

Chief Points of Chapter VII.

Principle of Archimedes.—When a body is submerged in a liquid, it loses weight equal to the weight of the liquid displaced by it.

Expressed differently, the upthrust experienced by an object in water is equal to the weight of the water displaced.

The number of cubic centimetres in an object is also the number of cubic centimetres of water displaced by such an object when it is immersed in water.

Flotation.—When a rod is floated in water the weight of the whole of it is equal to the weight of the liquid displaced by the portion of the rod immersed.

Applications of the Principle of Archimedes.—This principle, we have seen, explains why an iron ship can float and why balloons will rise in the air.

Density.—Equal volumes of different substances have different masses.

Absolute density is the mass of a unit volume of a substance.

Relative density or *specific gravity* is the ratio of the weight of any volume of a substance to the weight of an equal volume of water at temperatures other than 4° C.

$$\text{Relative density} = \frac{\text{weight of substance}}{\text{weight of equal volume of water}}.$$

$$\text{Relative density} = \frac{\text{weight of substance}}{\text{loss of weight in water}}.$$

Relative Density or Specific Gravity Bottle.—This is a small glass flask holding about 50 grams of water. It is provided with a nicely-fitting ground stopper, which is in the form of a tube with a very small hole through it. Its use depends upon the first of the preceding equations.

Balancing Columns of Liquids.—The densities of two liquids balanced in a U-tube are in inverse proportion to the lengths of the columns. Or, the density of one liquid is less than the density of the other in proportion as its column is longer than the balancing column of the other.

$$\therefore \text{ Relative density of a liquid} = \frac{\text{water column}}{\text{liquid column}}.$$

Hare's Apparatus.—Two straight glass tubes are connected at the top by a three-way junction, upon the unconnected end of which a piece of india-rubber tubing is placed. The lower ends of the tubes dip into beakers containing the liquids the relative densities of which have to be determined. By applying suction to the free end of the rubber tube the two liquids are drawn up the glass tubes. The heights of the liquid columns are inversely proportional to the relative densities of the liquids.

Relative Density of Gases.—(1) Find the weight of pure dry hydrogen which fills a flask under known conditions of temperature and pressure; (2) find the weight of an equal volume of gas under the same conditions. Then

$$\text{Relative density of gas} = \frac{\text{observed weight of gas}}{\text{observed weight of hydrogen}}.$$

QUESTIONS ON CHAPTER VII.

1. State the principle of Archimedes. Describe an experiment which illustrates the principle.

2. What determines the distance to which a rectangular wooden rod will sink when floated in a jar of water?

3. Describe an experiment to prove that when a rectangular block of wood is floated in water the weight of the water displaced by the portion immersed is equal to the weight of the whole block.

4. Give a few applications of the principle of Archimedes.

5. Carefully distinguish between the following terms:—density, absolute density, relative density.

6. Give a detailed description of the hydrostatic method of determining relative densities.

7. A glass stopper is weighed in air and then immersed successively in water, beer, milk, and vinegar, and the loss of weight noticed in each case. Explain how you would proceed to calculate the relative densities of each of the liquids from these observations.

8. How is the relative density of a liquid determined with a specific gravity bottle?

9. Explain a simple method for ascertaining the relative density of small shot or tin-tacks.

10. Being provided with two pieces of glass tube and a piece of india-rubber tubing, explain how you would proceed to (i) compare the relative densities of olive oil and spirits of wine, (ii) ascertain the relative density of a specimen of milk.

11. Describe the instrument known as Hare's apparatus. What advantages accrue from using it to determine the relative densities of liquids.

12. A piece of iron, weighing 275 grams, floats in mercury of density 13·59 with $\frac{5}{9}$ of its volume immersed. Determine the volume and density of the iron.

13. The specific gravity of brass referred to water is 8; taking the mass of 1 cubic foot of water as 1,000 ounces, find the density of brass in ounces to the cubic inch.

14. A cylinder whose base is a circle 1 foot in diameter, and whose height is 3 feet, weighs 10 lbs. Calculate its density, assuming 1 cubic foot of water to weigh 62·5 lbs.

15. Of two bodies one has a volume of 5 cubic inches, the other of one-fifth of a cubic foot; in a perfectly just balance the former weighs 15 oz., the latter 12·8 lbs. What is the ratio of the mass of the first to that of the second, and what is the ratio of the density of the first to that of the second.

16. If 100 cubic inches of oxygen (under certain circumstances of pressure and temperature) weigh 35 grains, and a cubic inch of

mercury weigh 49 lbs., how many cubic inches of oxygen would contain the same quantity of matter as a cubic inch of mercury?

17. In a certain state of the atmosphere 100 cubic inches of air weigh 31 grains; at a temperature the same as that of air 30 cubic inches of mercury weigh 14·88 lbs. Find the number of cubic inches of air which contain as much matter as a cubic inch of mercury.

18. If 28 cubic inches of water weigh a pound, what will be the specific gravity of a substance, 20 cubic inches of which weigh 3 lbs.

19. The volume of A is 3 cubic feet and its mass is 4 cwt.; the volume of B is 24 cubic inches and its mass is 35 oz. Find the ratio of the density of A to the density of B.

20. Describe a method of determining the specific gravity of a liquid.

21. Describe the method of determining the specific gravity of a body by weighing it alternately in air and water.

If a body weighs 14·4 grams in air and 12 grams in water, what is its specific gravity?

22. Explain what is meant by specific gravity. A body of specific gravity 5 weighs 20 grams in air; what will the body weigh when immersed in water.

23. How would you determine the volume of a pebble in cubic centimetres?

A number of nails are driven into a rough piece of wood, one cubic centimetre of which weighs 0·5 gram. It is required to find the weight of the nails without pulling them out. How could this be done?

24. A covered tin canister having a volume of 88 cubic centimetres contains just enough shot to sink it to the top of the cover when placed in cold water. Determine from this information

(*a*) The weight of the canister and shot.

(*b*) The weight of the water displaced by the canister.

25. Define mass, volume and density, and state the relation that exists between them.

Suppose you were given two irregular pieces of metal, one of which was gold and the other gilded brass. How would you find out by a physical method which piece was gold?

26. Explain why a ship made of iron will float in water, though iron itself is heavier, bulk for bulk, than water?

CHAPTER VIII

ATMOSPHERIC PRESSURE AND BOYLE'S LAW

SURROUNDING the earth in every latitude, over land and sea, is a gaseous envelope which is spoken of as the air or the atmosphere. Its presence when at rest is unperceived, though in motion it becomes apparent, since, by imparting its velocity to trees and other bodies free to move, it affords a demonstration of its existence. The student has already learnt to regard it as a form of matter, and as consequently possessing weight. The following experiments supply sufficient evidence of its existence and of its weight.

EXPT. 99.—Invert a so-called *empty* bottle under water in the pneumatic trough. Notice the bubbles which rise as the water flows into the bottle. The water displaces the air which thus becomes apparent.

EXPT. 100. Move quickly across the room with a drawing board in your hands. First hold the board "end on" and then "broadside on." Notice that in the first case little or no resistance is felt, while in the second, one's motion across the room is considerably impeded.

Weight of the Air.—It is easy to prove by direct experiment that the air has weight.

EXPT. 101.—Fit a one-holed india-rubber stopper into a fairly large glass flask, and fit into the stopper a short tube with a stop-cock upon it (Fig. 46). Put a little water in the flask; open the stop-cock; and boil the water. After boiling for a little time, turn off the tap and place the flask on one side to cool. When the flask is cool, weigh it, or counterpoise it. Then open the stop-cock; air will be heard to rush into the flask, and as it does so the balance will show an increase of weight.

If the air is completely removed from a flask by means of an air-pump, the difference in the weighings, before and after, will provide the exact weight of a given volume of air. Thus if the vessel has a capacity of a cubic foot the difference of weight will be found to be nearly an ounce and a quarter.

Pressure of the Atmosphere.—It has been seen that it is a property of all fluids that they communicate pressure in all directions, and consequently it is a character of air. It is a consequence of this fact that we are able to move about quite freely in spite of atmospheric pressure. Our bodies are subjected to an enormous pressure due to the whole weight of the atmosphere above us, and yet we are quite ignorant of it, at all events under ordinary circumstances. Why is this? The lungs which fill up a large part of our chest capacity are inflated with air, and other inside parts of the body are similarly in free communication with the atmosphere. The inside air is under just the same pressure as that outside, and consequently there is an exact compensation, and we are not crushed, as the student will perhaps have expected we ought to be.

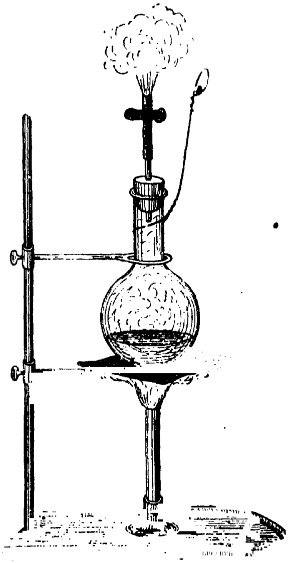

FIG. 46.—Experiment to show that Air has Weight.

EXPT. 102.—Procure a thin tin can having a neck, into which fits an india-rubber stopper. Take out the stopper and boil a little water in the can. After the water has been boiling for some time, so that the can is practically filled with steam, remove the can from the flame, and quickly put in the stopper as tightly as you can. After a few minutes the can will collapse.

The explanation of the effect produced in this experiment is that as the can cools the steam inside is condensed into water, and so occupies a much smaller volume. The pressure which the steam exerts on the inside of the can is thus removed, while the pressure of the air on the outside remains practically the same, the result being that the can is crushed. At the sea-level, under ordinary conditions, the pressure of the air is 15 lbs. on every square inch.

The following experiments also illustrate effects of atmospheric pressure :—

EXPT. 103.—Moisten a leathern sucker, press it upon a flat stone, and show that it can only be pulled off with difficulty, owing to the pressure of the atmosphere upon its upper surface.

EXPT. 104.—Dip the open end of a glass syringe or squirt into a bowl of water. Pull up the piston, and see that the water follows it, owing to the pressure of the atmosphere upon the surface of the water in the bowl. The action of a pump is very similar to this.

EXPT. 105.—Take a tumbler or cylinder with ground edges and completely fill it with water. Place a piece of stout writing paper across the top and invert the vessel. If the air has been carefully excluded from the cylinder the water does not run out (Fig. 47). Think what keeps the paper in its place.

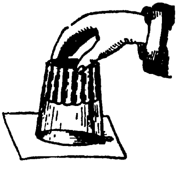

FIG. 47.—Experiment to illustrate the Pressure of the Atmosphere.

Principle of the Mercurial Barometer.—It has been seen that the air has weight, and that it exerts great pressure on the earth's surface ; we have now to learn how this pressure is measured.

EXPT. 106.—Procure a barometer tube and fit a short piece of india-rubber tubing upon its open end. Tie the free end of the tubing to a glass tube about six inches long open at both ends. Rest the barometer tube with its closed end downwards and pour mercury into it (being careful to remove all air bubbles) until the liquid reaches the short tube. Then fix the arrangement upright as in Fig. 48.

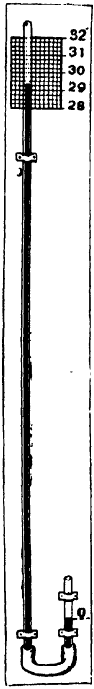

FIG. 48.—To explain the Principle of the Barometer.

The mercury in the long tube will be seen to fall so as to leave a space of a few inches between it and the closed end. The distance between the top of the mercury column in the closed tube and the surface of that in the open tube will be found to be about thirty inches.

The instrument used in Expt. 106 is evidently similar to a U-tube. Referring to Fig. 48, it is clear that there is a column of mercury supported by some means which is not at first apparent, or else the mercury would sink to the same level in the long and the short tube, for we know that liquids find their own level. If a hole were made in the closed end of the tube this would immediately happen. There will be no difficulty from what has been already said, in understanding that the column of mercury is kept in its position by the weight of the atmosphere pressing upon the surface of the mercury in the short open tube. The weight of the column of mercury and the weight of a column of the atmosphere with the same sectional area is exactly the same; both being measured from the level of the mercury in the short stem of the apparatus shown in Fig. 48, the mercury column to its upper limit in the long tube, the air to its upper limit, which, as will be seen, is a great distance from the surface of the earth. If for any reason the weight of the atmosphere becomes greater, the mercury will be pushed higher to preserve the balance; if it should become less, then similarly the amount of mercury which can be supported will be less, and so the height of the column of mercury is diminished.

The height must in every case be measured above the level of the mercury in the tube or cistern open to the atmosphere. In the arrangement shown in the accompanying illustration, a line is drawn at a fixed point O, and the short tube is shifted up or down until the top of the mercury in it is on a level with the line.

The student will now understand why it is so necessary to remove all the air bubbles in Expt. 106. If this were not done, when the tube was inverted the enclosed air would rise through the mercury and take up a position in the top of the tube above the mercury. The reading would not then be thirty inches, for instead of measuring the whole pressure of the atmosphere, what we should really be measuring would be the difference between the pressure of the whole atmosphere and that of the air enclosed in the tube. In a properly constructed barometer, therefore, there is nothing above the mercury in the tube except a little mercury vapour.

An arrangement like that described constitutes a **barometer,** which we can define as **an instrument for measuring the pressure exerted by the atmosphere.**

> EXPT. 107.—The preceding experiment will have shown you that air pressing upon the surface of the mercury in the short open arm of the U-tube will balance a long column of mercury in the closed arm. Slip a piece of india-rubber tubing upon the open end and notice what happens when you blow sharply into it. Suck air out of the tube, and observe the result.

These experiments show you the effect of increasing and decreasing the pressure upon the free surface of the mercury. How will the height of the mercury column be effected (1) if the pressure of the air decreases, (2) if the pressure of the air increases?

Weight of Column of Atmosphere.—The following is another form of the experiment to show atmospheric pressure by means of a barometer.

> EXPT. 108.—Procure a thick glass tube about 36 inches long and closed at one end. Fill the tube with mercury; place your thumb over the open end; invert the tube; place the open end in a cup of mercury and take away your thumb.

A column of mercury will be supported in the tube by the pressure of the atmosphere. The distance between the top of the column and the surface of the mercury in the cup will be about 30 inches, or 76 cm. when the tube is vertical (Fig. 49, *b*). If the tube is inclined so that the closed end of it is less than this height above the mercury in the cup (Fig. 49, *c*) the mercury fills it completely; and if the tube is less than 30 inches long, it is also filled by the mercury (Fig. 49, *a*). On an average the atmosphere at sea-level will balance a column of mercury 30 inches in length. No matter if the closed tube is 30 feet long, the top of the mercury column will only be about 30 inches above the level of the mercury in the cistern.

EXPT. 109.—Weigh the column of mercury sustained in the tube and measure the diameter of the tube. The area of the bore can then be found (area of circle = $\text{radius}^2 \times 3\frac{1}{7}$). Calculate from these observations the weight of a column of air on any given area.

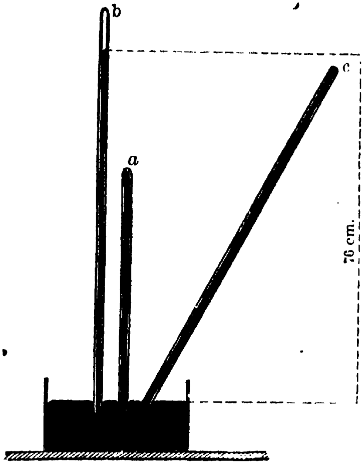

FIG. 49.—Height of Mercury in Barometer.

If the tube had a bore with a sectional area of exactly one square-inch, there would be 30 cubic inches of mercury in a column 30 inches long; and since a cubic inch of mercury weighs about half a pound, the whole column would weigh 15 lbs. This column balances a column of air of the same area, so that we find that the weight of a column of air upon an area of one inch is 15 lbs. when the barometer stands at 30 inches.

Mercury a Convenient Liquid for Barometers.—The use of mercury for barometers is a matter of convenience. Since the column of mercury which the atmosphere is able to support is 30 inches high, it is clear that, as water, *e.g.*, is 13·6 times as light as mercury, the column of water which could be supported would be 30 × 13·6 = 408 inches = 34 feet, which would not be a convenient length for a barometer. The length of the column of glycerine which can be similarly supported is 27 feet. But in the case of lighter liquids like these, any small variation in the weight of the atmosphere is accompanied with a much greater alteration in the level of the column of liquid, and in consequence it is possible to measure such variations with much greater accuracy. For this reason barometers are sometimes made of glycerine. There is one in the Western Galleries of the South Kensington Museum, which the student could see at any time.

Pressure of the Atmosphere at Different Altitudes.—The atmosphere being a material substance, the longer the column of it there is above the barometer, the greater will be the weight of that column, and the greater the pressure it will exert upon the mercury in the barometer. Hence, as we ascend through the atmosphere with a barometer, we reduce the amount of air above it pressing down upon it, and in consequence the column of mercury the air is able to support will be less and less as we ascend. On the contrary, if we can descend from any position, *e.g.*, down the shaft of a mine, the mercury column will be pushed higher and higher as we gradually increase the length of the column of air above it. Since the height of the column of mercury varies thus with the position of the barometer, it is clear that the variation in its readings supplies a ready means of ascertaining the height of the place of observation above the sea-level, provided we know the rate at which the height of the barometer varies with an alteration in the altitude of the place. The rule which expresses this relation is not a simple one, but for small elevations it is said that a rise or fall of one inch in the height of the barometer corresponds to an alteration of 900 feet in the altitude of the barometer.

Relation between Volume and Pressure of a Gas.—Boyle's Law.—Before we can properly understand how and why the density of the atmosphere varies, it is necessary to become acquainted with the rule expressing the relation between the volume and pressure of a gas. A convenient and efficient arrangement for determining this relation is represented in Fig. 50, and it has the merit of being easily constructed with the materials which should be found in every physical laboratory. A burette, A, is fixed to an upright stand by means of wire or small brass clamps. Upon the lower end a piece of thick india-rubber tubing is firmly tied. The tubing is also tied upon one end of a three-way piece, B. A long glass tube, C, connected with another end of the three-way junction is supported upon the stand side by side with the burette. To the third end of the junction a long piece of rubber tubing, D, is fixed, and at the top of this a large funnel or bottle, E, with the bottom cut off, is tied. The use of this apparatus will be seen by the following experiments:—

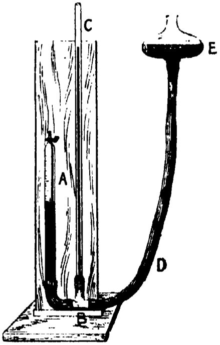

FIG. 50.—Apparatus to show the Relation between the volume and Pressure of a Gas. (Boyle's Law).

EXPT. 110.—Open the stop-cock of the burette. Pour mercury into the bottle until its level is about half-way up the burette. This mercury is evidently at the same level in the burette, long glass tube, and rubber tube, and the air in the burette is at the pressure of the atmosphere. Now close the stop-cock. We have enclosed a given volume of air, which is indicated by the length of tube it occupies. Pour mercury into the bottle until the bottle is nearly half full. By lifting or lowering the bottle the level of the mercury is, of course, varied. Move the bottle until the difference between the level of the mercury in the burette and in the straight glass tube is 30 inches, or whatever is the reading of the barometer at the time of the experiment.

Measure the length of the air column which is now subjected to a pressure equal to two atmospheres, and notice that it is one-half of the original volume. That is, by doubling the pressure we have halved the volume.

Repeat the experiment with the bottle at any height, and when the apparatus has cooled down measure the difference in the levels of the mercury in the two tubes and also the length of the air column.

Boyle's Law.—If we multiply the total pressure (obtained by adding the height of the barometer to the difference in level of the mercury in the burette and the open tube) by the volume of the air we shall notice that the result is always the same. This relation was discovered by Boyle, and is known as **Boyle's Law.** It can be expressed by saying **that when the temperature remains the same, the volume of a gas varies inversely as its pressure.** Or, what is the same thing, **the temperature remaining the same, the product of the pressure into the volume is constant.**

EXPT. 111.—Perform several experiments with the Boyle's Law apparatus, observing in each case (*a*) the volume occupied by the air in the burette, (*b*) the difference in level between the mercury in the closed burette and the open tube, or "head" of mercury. Tabulate your results thus:—

"Head" of Mercury.	Height of Barometer.	Total Pressure (P).	Volume of Air in Burette (V).	(P × V).

When the air occupies less volume than it did before the stop-cock was closed it is evidently under a pressure greater than that of the atmosphere; when it occupies a greater volume it is under a pressure less than that of the atmosphere. In either case, however, the pressure to which the air is subjected multiplied by the volume it occupies gives a constant product. But in performing the above experiment at pressures less than that of the atmosphere, the difference between the heights of the mercury in the open and closed tubes must be *subtracted* from the barometric height and not added to it. Why?

Application of Boyle's Law to the Atmosphere.—It has been learnt that if we increase the volume occupied by a given mass of a substance we decrease its density, or if we decrease its volume we increase its density. Therefore, we not only alter the volume of the enclosed air in the two preceding experiments, but we also change its density. Increase of density and increase of pressure are proportional to one another. It is not difficult to apply these facts to the case of the atmosphere. We have learnt that the pressure of the atmosphere decreases as we ascend, and we are now able to add that its density decreases also and at the same rate. Therefore the densest atmosphere will be that at the surface of the earth, leaving out, of course, the air of mines and other cavities below the surface, where the air will be denser still. The air gets less dense or rarer as we leave the surface, until eventually it becomes so rare that its existence is practically not discernible.

CHIEF POINTS OF CHAPTER VIII.

Determination of the Weight of Air.—This is done by weighing a stoppered flask (1) full of air, (2) after the air has been driven out. The difference of weight is the weight of the air which filled the flask.

Pressure of Air.—In consequence of its weight, air exerts pressure. On the earth's surface, under ordinary conditions, the pressure of the atmosphere equals 15 lbs. per square inch.

A Barometer is an instrument used to measure the pressure of the atmosphere.

The Principle of the Action of a Mercurial Barometer.—The column of mercury balances a column of air extending from the surface of the mercury in the cup to the limits of the atmosphere. The action is therefore analogous to balancing columns of different liquids in a U-tube.

Some Points referring to Barometers.—(1) Mercury is the liquid usually employed because it is the heaviest liquid known, is not very volatile, is easy to see, and does not wet the tube. (2) Water and other liquids can be used in the construction of barometers, but these barometers require to be longer in order to hold the longer liquid columns required to balance the atmospheric pressure. (3) If a crack or a hole is made in the top of a barometer, air enters through it, and the mercury column falls to the level of the liquid in the cup. (4) The height of the column of mercury is about 30 inches at sea-level, but it varies from time to time on account of variations of atmospheric pressure.

Pressure of the Atmosphere at Different Altitudes.—

Altitude.	*Height of Barometer.*	*Pressure in lbs. per sq. in.*
Sea-level	30 inches	15
3½ miles	15 ,,	7½

Boyle's Law.—When the temperature remains the same, the volume of a gas varies inversely as its pressure.

Or, the temperature remaining the same, the product of the pressure into the volume is constant.

Boyle's Law applied to the Atmosphere.—When the temperature remains the same, the volume of a gas varies inversely as the pressure. If a cubic foot of air at sea-level, where the pressure is 15 lbs. per square inch, were taken up to a height of 3½ miles, where the pressure is one-half, viz., 7½ lbs. per square inch, it would expand to two cubic feet.

QUESTIONS ON CHAPTER VIII.

1. How can it be proved that the air has weight? Illustrate your answer by a few simple experiments.

2. Given a glass tube 32 inches long, closed at one end, a bottle of mercury (quicksilver), and a small cup. State how you proceed (*a*) to construct a barometer, and (*b*) to show the readings of this barometer.

3. (*a*) State the average height of the mercury in a barometer at the sea-level and at the top of a mountain 3½ miles high.
   (*b*) What is the cause of the difference in the height of the mercury column?
   (*c*) What do you know concerning the condition of the upper layers of the atmosphere?

4. Describe the construction and use of an ordinary mercurial barometer.

5. How can the weight of air be determined? In what way is the pressure exercised by the atmosphere on the earth's surface in consequence of its weight, stated? How is it that we are able to move about under the weight of the atmosphere?

6. Explain the chief reasons why mercury is the liquid usually employed in the construction of barometers.

7. State the principle on which the action of a mercurial barometer depends Why is a water barometer longer than a mercurial barometer? What occupies the space above the mercurial column in the latter instrument? If a hole were bored through the glass above the column of mercury, what would happen?

8. (*a*) Why does the mercury stand higher in the tube than in the cup of a barometer?
   (*b*) What is the average height of the mercury in a barometer tube at the sea-level?
   (*c*) Why does the height vary from time to time?
   (*d*) Why is the barometer regarded as a "weather glass"?

9. If you ascended to the height of 3½ miles in a balloon, carrying a barometer, state—

(*a*) The indication which would be given by the barometer.

(*b*) Your explanation of this.

10. State Boyle's Law. Describe an experiment which demonstrates the truth of the law.

11. What volume would the quantity of air which fills a space of a cubic foot at the sea-level occupy if carried to a height of 3½ miles? Explain this alteration in volume.

12. What would be the effect on the height of the mercury column if an ordinary barometer were carried down a deep mine?

CHAPTER IX

WORK AND ENERGY

Work.—Newton's first law of motion teaches us that a body at rest is only set in motion by the action of a force upon it, and also that a moving body only changes the direction of its motion, or its speed, as the result of the action of a force. When a body moves from rest the continued action of the force upon it causes an acceleration in the body.

In the case of a body already moving, though we have every right to argue that a change of direction or a change of velocity is the result of an external force, we cannot apply the converse statement and say that an external force acting upon a moving body causes a change of direction or of velocity, for in some instances the force may be entirely occupied in maintaining such motion in opposition to other forces acting upon it. Thus, when a ship is sailing with a uniform speed, the force of the wind is exhausted in maintaining this velocity by overcoming the resistance of the water.

When a force acts in either of these ways it is said to do **work,** that is, work is done by a force in setting a body at rest into motion and giving it a regularly increasing velocity, or by maintaining a uniform motion in opposition to the action of other forces. We may class all these forces acting in opposition to the force which is being considered under the inclusive name of *resistance*.

We shall thus obtain for our definition of work the following statement:—**Work is done by a force, either when it acts upon a body producing an acceleration in its velocity, or when it maintains a uniform velocity in a body in opposition to resistance.** Or, more briefly, **Work is done when the point of application of a force moves.** A little consideration will convince the student that all instances of mechanical contrivances, by means of which work is accomplished, come within the scope of our definition, as well as all other cases when we say in ordinary language that work is done. Take, for instance, a horse drawing a heavy weight along a road. Here the force exerted by the horse is used up in overcoming the resistance due to the road. A man raising a mass from the ground overcomes the resistance due to the body's weight. A body falling from a height under the influence of the earth's attractive force has work done upon it, with the result that its velocity increases according to a uniform acceleration of 32·2 feet per second in every second.

Measurement of Work.—Referring to our definition of work it would seem as though we had two kinds of work to measure, viz., the *work of acceleration* and the *work against resistance.* But since we can make a force perform either of these kinds of work according to the condition under which it acts, it is possible to measure either of them in the same units. An example will make this clearer. We can either allow a mass to drop from the hand and to move freely through the air with the uniform acceleration we have mentioned above until it reaches the ground; or we can attach the mass to a string, pass the string over a cylinder, and allow it to move towards the earth with a small uniform velocity—a result which can be brought about by applying the necessary friction between the cord and the cylinder, that is by applying a resistance. The final result brought about is the same under both sets of conditions; but in the first case the work is of acceleration, while in the second it is work against resistance.

Unless there is motion no work is done. If we put a weight upon a table or shelf, so long as the weight remains in one place, it evidently does no work, though it is capable of doing work by reason of its elevated position.

For practical purposes the unit of work which is adopted is *the work done in raising the mass of one pound through one foot*, and it is called the **foot-pound.** This is not a strictly constant unit, for it will be evident, in the light of what has been said about the weight of a body, that where the weight is greater the amount of work done will be greater. The unit of work will vary slightly in different latitudes in a precisely similar manner to that in which the weight of a mass varies.

It will have been noticed that the question of time does not enter into an estimation of the amount of work done. It is manifest that the same quantity of work is accomplished whether a day is spent in raising a weight to a given height from the ground or only a minute. If we introduce the time taken to perform the work we begin to consider what is called the **power** of the agent. We should measure this power by the quantity of work the agent can perform in a given time ; or **power is the rate of doing work** and is measured by the work done in a second. Thus, engineers use the expression **horse-power,** by which they mean the rate at which a good horse works. James Watt estimated this at 33,000 foot-pounds per minute, or 550 foot-pounds a second.

Generally, then, to find the amount of work performed by any force, we multiply the value of the force (expressed in suitable units) by the space through which it acts (using the corresponding unit in measuring this quantity also). The following simple application of the rule will familiarise the student with the method for employing it :–

How much work is done when an engine weighing 12 tons moves a mile on a horizontal road, if the total resistance is equal to a retarding force of 10 lbs. weight per ton ?

The total resistance equals $12 \times 10 = 120$ lbs. weight, the distance traversed is 5,280 feet.

$$\therefore \text{ Work done} = (120 \times 5{,}280) \text{ foot-pounds.}$$

Energy.—By the energy of a body we mean its power of overcoming resistance or doing work. All moving bodies possess energy. Moving air or wind drives round the sails of a windmill and so works the machinery to which the sails are attached ; it drives along a ship, thus overcoming the resistance of the water. The running stream works the mill-wheel and the energy it possessed is expended in grinding corn. The bullet fired from a rifle can pierce a sheet of metal by overcoming the cohesion between its particles.

Expt. 112.—Stretch a piece of tissue-paper over the top of an empty jam-pot. Carefully place a bullet on the paper and notice the paper will support it. Now lift the bullet and allow it to drop on to the paper. It is seen that the bullet pierces the paper.

Expt. 113.—Support a weight by a thin thread. Show that though the thread will support the weight at rest it will be broken if the weight is allowed to fall.

EXPT. 114.—Show that a falling weight attached by a string to a spring balance extends the balance beyond the point which it indicates when the weight is at rest.

All these examples are cases of the energy of moving bodies, or the energy of motion, or *Kinetic Energy*. **Kinetic Energy is the energy of matter in motion. All energy which is not kinetic is known as Potential Energy.** It is capable of becoming kinetic, or active when the conditions become suitable. Imagine a mass raised from the ground and placed upon a high shelf. We know that to place it in this position we must expend a certain amount of work, which is measured by multiplying its weight by the height through which its mass is raised. Further, we know that just as soon as we release it from its position of rest, making it free to move, it will travel with an ever-increasing velocity until it reaches the ground. On the shelf the mass, by virtue of its position, possessed a certain amount of potential energy exactly equal to the work expended in placing it there.

Similarly, an ordinary dining-room clock, which is worked by a spring, affords us an example of potential energy. The wound-up spring possesses potential energy exactly equal to the amount of work done in winding it up. This potential energy is being continually converted into kinetic energy as it becomes unwound in working the clock.

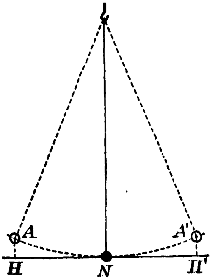

FIG. 51.—Pendulum in Oscillation.

The motion of a pendulum affords an interesting example of the two forms of energy. At the end of its swing, in the position A (Fig. 51), the bob of the pendulum possesses potential energy enough to carry it through half an oscillation,[1] that is, until it reaches its lowest position N, when the whole of the energy of position which it possessed at A is expended, as it can reach no lower position. But though it lacks potential energy, since it is a mass moving with the velocity it has gained in its passage from A to N, it possesses energy of motion or kinetic energy enough to carry it up to its next position of rest at A′—where the only energy it will have will be again potential. Through the next oscillation from A′ to A it will pass through just the same transformations again.

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[1] Some physicists regard the motion from A to A^1 as half an oscillation.

At any point in the swing the pendulum will possess a certain amount of energy due to position and a certain amount due to motion, but the total amount of energy—the sum of the potential and the kinetic energy—is always the same, the loss of one form of energy being exactly equalised by the gain of the other.

Measurement of Kinetic Energy.—When we wish to measure the energy of moving bodies we have to find an expression which will be equal to the amount of work these bodies are capable of performing when the whole of their energy is converted into work. Such an expression is easily found from first principles, as the reader will learn as he pursues his studies.

If M represents the mass of such a moving body and v its velocity, the expression $\frac{Mv^2}{2}$ is a measure of its energy, and gives us a means of calculating the energy possessed by any body in motion in terms of its mass and its velocity. If we wish to express it in foot-pounds we shall, as we have seen, divide its value by g.

Kinetic energy, or the energy of moving bodies, is equal to one-half the product of the body's mass and the square of its velocity. Hence, the measure of energy is dependent on the units of mass and velocity employed.

Forms of Energy.—A body may possess energy due to other causes than that of the actual motion of the body as a whole. When it is in rapid vibration, or when it is heated, or when it is electrified, it is endowed with energy in consequence of these conditions. But when a body is in rapid vibration it gives out sound or becomes a sounding body, hence we may regard sound as a form of energy. We shall see that work may be done by the passage of heat from a hot body to a cold one, and, in consequence, heat is properly regarded as another form of energy. An intensely hot body emits light, hence it would seem that light and heat have a common cause and that we must also regard light, like heat, as a manifestation of energy. When a body is electrified it has the power of attracting unelectrified and certain electrified bodies also, and when such bodies are attracted as a result of this electrification we see that electrification must similarly be looked upon as still another kind of energy. But it must be borne in mind that electrification is not electricity. Then, too, there is the attraction of magnetism, which is capable of accomplishing work, and hence must likewise be looked upon as a form of energy. Chemical combinations, again, are always accompanied by the development of heat, and resulting as they do from the chemical attraction of two more or less unlike forms of matter, we shall be right in saying chemical combination are always accompanied by energy changes and so in regarding chemical attraction as another kind of energy. **In addition to the energy of moving bodies we have energy manifested as sound, heat, light, electrification, magnetism, and chemical action.**

Bearing this in mind, it will perhaps assist the student to grasp the enlarged conception of energy which is here presented to him, if he regards *energy as being a capacity for producing physical change.*

Heat as a Form of Energy.—Heat was not always regarded in this way. It was originally thought to be a fluid called Caloric, and it was supposed that a piece of hot iron differed from a cold piece in having entered into some sort of union with this fluid. But, since the experiments of Rumford, we can no longer doubt that heat is not material, but a form of energy. Rumford boiled water by the heat developed by the friction between two metal surfaces which he rubbed together; and he found that the amount of water he could bring to the boiling temperature depended only on the amount of work he expended in rubbing. Since he could obtain an indefinite amount of heat from two definite masses of metal, it was quite clear that heat could not be matter, which, as we have seen, cannot be created. Davy made the truth even clearer by obtaining heat enough to melt ice by simply rubbing two pieces of this solid together. They were both cold or without caloric; and since heat could be obtained by rubbing even these together, it was quite certain that heat could not be a fluid. Joule went a step further and measured the amount of work which must be done to obtain a given quantity of heat; or, as we say, he measured the **mechanical equivalent of heat.**

Some examples which will be familiar to the student will provide him with proofs of the statement that heat and work are convertible. When a brake is applied to the wheels of a train, as it stops at a station, it is a common thing to see sparks fly. The resistance of friction which overcomes the motion of the train causes a sufficient amount of heat to be developed to raise the particles of steel, which get rubbed off, to a red heat. By continually hammering a piece of iron on an anvil it can be made too hot to hold in the hand.

The following experiments show that heat appears when motion is destroyed:—

EXPT. 115.—Procure a piece of lead in the form of a sphere (about the size of a marble) with a neck or hook upon which a piece of string can be fastened. Tie a piece of string firmly to the neck, and while holding the string strike the sphere several times smartly on an iron plate. Test the temperature of the ball before and after the experiment by a thermopile and galvanometer.[1]

EXPT. 116.—Hammer a piece of lead, or saw wood, and test the temperature of the lead or saw before and after the experiment.

EXPT. 117.—Rub a brass nail or button on a wooden seat, and notice its increase of temperature.

When we rub a lucifer match along a rough surface the heat into which the work is converted is enough to ignite the match. In all these cases mechanical work is converted into heat. The converse is true also, heat is convertible into work. In the steam-engine the heat of the furnace changes the water in the boiler into steam. The steam forces the piston along the cylinder, and this movement of the piston in a straight line is converted into the circular motion of a fly-wheel; or is used, through the intervention of suitable mechanism, in pumping water or performing some other kind of work. The steam which enters the cylinder is hotter than that which leaves it for the condenser. Thus, we see, part of the heat of the steam has been converted into useful work and parts of it have been lost to the condenser, the air, etc.

We can show by a simple experiment that heat often disappears when motion is produced.

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[1] At this stage the thermopile and galvanometer must simply be regarded as a delicate means of measuring temperature.

Expt. 118.—Allow air, which having been compressed into a cylinder has again assumed the temperature of the air, to come in contact with a thermopile or other delicate means of measuring changes of temperature. Notice the cooling of the compressed air when allowed to escape.

The Energy of Radiant Heat and Light.—We shall learn more fully later the reasons we have for regarding these as being of the same nature. We refer under this heading to those forms of energy which travel through space where, as we know, there is no air to convey them in the way in which sound is transmitted. But in order to understand the transmission of these forms of energy it is necessary to imagine the existence of a medium which is referred to as the luminiferous ether, or more shortly as the "ether." Certain experiments and observations by different investigators leave no doubt of the existence of this medium, vibrations in which cause light and radiant heat. Sound passes through air by the to and fro vibration, in turn, of the air particles, in the form of a *wave* as it is called. So radiant heat and light pass through the ether by the successive motions of the constituent molecules of the ethereal medium.

This takes place with astonishing quickness, for light travels about 186,000 miles per second, or something like 7½ times round the earth in this small interval of time. Nor does the ether fill inter-steller space alone, for it must exist in the interstices (p. 3.) of those bodies through which radiant heat and light can pass, or how else can light pass through a transparent body, or radiant heat through substances like rock-salt?

Radiation can be converted into work, but in a less direct manner than is the case with ordinary heat. It must first be absorbed and heat some material body causing its molecules to oscillate in the manner we have described. This form of heat, we have seen, has a mechanical equivalent, and we can fairly argue that, if the whole radiation is absorbed, the mechanical equivalent of the absorbed heat is an exact measure of the energy of the radiation.

Energy of Electrification.—We are careful not to speak of electricity as a form of energy, for whatever electricity may be it certainly is not energy. Though it would be very interesting to discuss the nature of electricity it does not come within the scope of our subject. The following experiment will show that we are right in regarding electrification as a manifestation of energy :—

EXPT. 119.—Suspend a pith-ball by a silk thread to a bent wire, as is shown in Fig. 52. Rub a rod of sealing-wax with fur and touch the pith-ball with the rod. Notice that after the contact it is impossible to make the ball come to the rod. They *repel* one another.

EXPT. 120.—Having touched the pith-ball with the rod of sealing-wax, which has been rubbed with fur as in the last experiment, bring near to it a rod of glass which has been rubbed with dry silk. Notice that the ball is *attracted* towards the glass rod.

What is the significance of these experiments? In both cases the pith-ball moves through a certain distance under the influence of a force, in one case of repulsion, in the other of attraction, and in consequence work is done. Under certain circumstances, as in the discharge of a Leyden jar, the energy of electrification becomes manifest in the form of a vivid spark and a slight explosive sound.

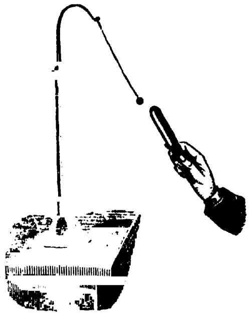

FIG. 52.—Electrical Attraction.

Electricity in motion constitutes what is known as the **electric current,** and of its capability of doing work the student has abundant evidence in the heat and light of an incandescent electric lamp, where the passage of the current through a wire, which offers considerable resistance to its passage, causes the wire to become sufficiently hot to be utilised as a source of light.

It will be very instructive to consider briefly the case where the electric current is formed as the result of chemical action in a battery, and thence passed by wires to a lamp of the kind mentioned. This is the ordinary condition of things as already described; but imagine the lamp left out and the battery made to work simply through an ordinary copper wire which presents little resistance. The current has no work to do beyond heating the wire, and the energy of the current is almost wholly expended in heating the liquids and other parts of the battery, which is of course a very undesirable waste of energy.

EXPT. 121.—Show the motion of a magnet produced by an electric current in a wire held over it.

Instances of the conversion of the energy of the electric current into mechanical work will doubtless have come under the student's attention.

Other forms of energy can be converted into that of electrification and of electricity in motion. If we heat certain crystals, *e.g.*, tourmaline crystals, it is found that they become electrified, one part of the crystal exhibiting electrification of the kind developed when sealing-wax is rubbed with fur, another exhibiting the kind obtained by rubbing glass with silk.

Heat can give rise to electric currents. If we solder a piece of the metal antimony to a piece of the metal bismuth and apply heat to the junction, it is found that an electric current passes from the bismuth to the antimony.

Energy of Chemical Action. EXPT. 122.—Place a small piece of dry phosphorus[1] on a plate, and a short distance from it a few grains of solid iodine. Nothing happens. By means of a glass rod push the piece of phosphorus on to the iodine, and notice that when they come into contact the phosphorus inflames and dense fumes are formed which, as will be understood after reading the chemical section of this book, are a compound of phosphorus and iodine.

EXPT. 123.—Call attention to the heat of the flame of the laboratory burner, where certain chemical actions are going on, which are described in a later chapter, (Chapter XV.)

In speaking of heat as a form of energy we took the example of the work done by an engine as the result of the heat from the furnace; but we can now push our inquiry a step further back. What causes the heat of the furnace? Evidently the burning of the coal, which, as we shall learn, is nothing more than chemical action. The coal enters into a chemical combination with one of the constituents of the atmosphere, and in doing so, heat is evolved, as it is, indeed, in all cases of chemical combination.

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[1] Great care must be used in handling phosphorus, as it is easily ignited.

But as we shall soon have occasion to study many cases of chemical action, we need not spend much space over the matter here, though it must be pointed out that, just as in all the other instances of energy we have studied, not only is it true that chemical action is accompanied by changes of energy, but also that it can result from these other forms. Light can be made to bring about chemical action, as it does in the case of the exposed photographic plate. Electric separation or the energy of electrification also causes chemical action, as will be seen later.

Transformation of Energy.—We have learnt that one kind of energy can cease to exist in that particular form, and can assume another condition. We have seen that the energy of moving bodies can give rise to sound and heat; that heat can be changed into the energy of moving bodies, electric currents, and chemical action. Indeed, one form of energy can assume almost any other form. The general tendency of all forms of energy is gradually to get converted into heat. When this change has become complete and all the energy of the universe exists as heat at the same temperature, there will be no further transformations possible. Consequently no work of any kind will be possible, which means there will be no life, no movement of bodies from place to place—a still, dead world, in fact.

Conservation of Energy.—We have seen that matter cannot be destroyed; we have now to learn that **energy is indestructible.** The total amount of energy in the universe remains the same. One form may be changed into another, but we can create no new energy. We may be unable to trace and account for some of it in the numerous transformations which it undergoes, but we are sure, from many considerations, that if our methods of experiment were only refined enough, we should be able to account for the whole amount.

The great source of energy in the solar system is the sun. It is from the sun that we are continuously receiving streams of energy in the form of radiation, which are continually assuming the various other forms of energy we have considered. Returning once more to our steam engine, we have traced back the work it does to the chemical combination of the coal with a part of the atmosphere. Or putting the case in another way, we have seen that there exists a certain amount of potential energy in the coal and atmosphere which is capable of becoming kinetic as soon as the temperature at which the coal can combine with oxygen has been reached. Whence comes the potential energy of the coal? Coal results from the compression of vegetable material which lived and flourished on the earth ages and ages ago. This vegetable material formed the tissues of mosses and other similar plants, which in the presence of sunlight have the power, by virtue of the green colouring matter they contain, of decomposing one of the gases of the atmosphere, carbon dioxide, splitting it up into its elements, carbon and oxygen, reserving the former for themselves and returning the latter to the air.[1]

This carbon unites with the elements of water contained in the plants forming compounds which build up the tissues of which the plant is constructed. The tissues of the plant represent from our point of view the work done in splitting up the carbon dioxide by absorbing the energy of radiation. They still represent this energy when they have assumed the condition of the coal, and it is in this sense that coal is poetically referred to as "bottled sunshine."

If the earth receives so great an amount of energy from the sun, it is easy to understand that the total quantity of energy which is given out by the sun must be enormously greater. But this radiation from the sun's surface is continually going on; that is, the sun is constantly losing energy, and this cannot go on indefinitely without the loss being made good. How is the energy of the sun maintained? It has been suggested that the heat generated by the impact of the meteorites which fall upon the sun in great numbers is capable of accounting for this energy; and that in addition to this a slight shrinkage of the sun's mass in cooling evolves a large amount of energy. But interesting as this subject is we cannot pursue it further here.

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[1] The student must remember to re-read this after studying Chapter XV.

CHIEF POINTS OF CHAPTER IX.

Work is the act of overcoming resistance, or causing change of velocity.

The Principle of Work.—Work is done by a force when the point of application moves through a certain distance.

Work done = force × distance.

A Foot-pound is the work done when a force equal to the weight of one pound is exerted through a distance of one foot in the direction of the force.

Power, or the rate of doing work, is measured by the number of foot pounds produced in a given time. A *horse-power* is equal to 33,000 foot-pounds per minute.

Energy is the ability to do work.

Energy.

Kinetic,	Potential,
due to motion as shown by a moving cannon-ball and a flowing river.	due to position, as a mass, at the top of a tower and a wound-up spring.

Manifestations of Energy.

- *Sound,* produced by vibrations of material substances.
- *Heat,* produced by motions of molecules.
- *Radiation,* due to vibration set up in the ether, and producing the effects of radiant light, heat, &c.
- *Electrification,* evidenced by the action of electrified bodies upon one another.
- *Electricity in motion,* as in the electric current, which can be transformed into light, heat, or motion.
- *Chemical action,* used to produce light and heat when a fire burns, and to produce motion in a gas-engine.

Measurement of Kinetic Energy.—Kinetic energy is equal to one-half the product of the body's mass and the square of its velocity. But the mass and the velocity must be expressed in suitable units.

Conservation of Energy.—Energy is never lost, but only changed in form, and whatever transformations take place, the sum total of kinetic energy and potential energy remains the same.

Questions on Chapter IX.

1. In the case of a shot fired at a target, state (*a*) why the velocity of the shot changes; and (*b*) why the target is made hot where the shot strikes it.

2. Define work, and describe an experiment to prove that a falling ball is capable of doing work.

3. Explain the terms power of an agent and horse-power.

4. What is meant by a foot-pound of work? What is the value of a horse-power in terms of this unit?

5. How is kinetic energy measured? If we wish to express the result in foot-pounds, how do we proceed?

6. A man weighing 140 lbs. puts a load of 100 lbs. on his back and carries it up a ladder to a height of 50 feet. How many foot-pounds of work does he do altogether and what part of his work is done usefully?

7. A body weighing 10 lbs. is placed on a horizontal plane and is made to slide over a distance of 50 feet by a force of 4 lbs. What number of units of work is done by the force?

8. If a man can work at the rate of 210,000 foot-pounds an hour, how long would it take him to raise a weight of 10 tons through 150 feet, supposing him to be provided with a suitable machine?

9. A horse pulling a horizontal trace with a force equal to the weight of 72 lbs., draws a cart along a level road at the rate of 3¾ miles per hour. What amount of work is done by the horse in 5 minutes?

10. If a force equal to the weight of 10 lbs. revolve three times tangentially round a circle of 5 feet radius, find the work it would do.

11. A cannon-ball whose mass is 60 lbs. falls through a vertical height of 400 feet. What is its energy at the end of its fall?

12. What is the kinetic energy of a mass of 5 lbs. moving with a velocity of 10 feet per second? State clearly what the unit is in terms of which your answer is expressed.

13. A body whose mass is 10 lbs. is carried up to the top of a house 30 feet high. By how many foot-pounds has the change of position increased its potential energy? If it is allowed to fall, what number of foot-pounds of kinetic energy will it have when it reaches the ground?

14. A particle moving from rest is acted on through 250 feet by a force of 9 pounds. Find its kinetic energy; and its mass being 5 lbs., find its velocity.

15. What is the difference between kinetic energy and potential energy?

16. Describe an experiment to prove that energy due to visible motion can be transferred from one body to another.

17. What proof can you adduce that the energy of visible motion can be transformed into heat?

18. How do you account for the fact that heat generally appears when motion is lost?

19. Describe an experiment which indicates that "Heat is a form of Energy."

20. How would you prove that "when bodies are electrified they are endowed with energy and can produce motion and heat"?

21. Name some forms of energy, and give instances of the transformation of energy from one form to another.

22. How would you prove that the destruction of the visible motion of a body is often accompanied by the appearance of heat?

23. When a brake is applied to a wheel of a train red-hot sparks are seen. What are these sparks, what is the source of their heat and why do they soon disappear?

24. Energy may be defined as the ability to do work. Give instances to show that heat and chemical action are forms of energy thus described.

25. Define the meaning of the word "energy," and give three or four examples to prove that a moving body possesses energy.

CHAPTER X

HEAT

Heat.—In regarding heat as a form of energy we found it necessary to speak in somewhat vague terms of its nature, but the student will, by studying the properties of heated bodies, and by learning how it is measured, be in a position to form a much clearer conception of what constitutes heat.

Hot and Cold Bodies.—EXPT. 124.—Arrange three basins in a row, into the first put water as hot as the hand can bear, into the second put luke-warm water, and fill the third with cold water. Place the right hand into the cold water, and the left into the hot, and after half a minute put both quickly into the luke-warm water. Notice that the left hand feels cold and the right warm while in the same water.

It will be evident from this experiment that the sense of touch is not to be depended upon for accurately estimating the heat condition of a body. To be able to make an estimate of the temperature, or intensity of heat, of a body we must utilise some effect produced by heat upon an inanimate substance.

Effects of Heat.—The effects of heating a substance may be classed under three heads:—

(1) Change of size.
(2) Change of temperature.
(3) Change of state.

1. The change of size which a body undergoes is spoken of as the amount it *expands*; or heat is said to cause *expansion* in the body. This expansion is regarded in three ways. When we are dealing with solids, we speak of expansion in length or linear

expansion, expansion in area or superficial expansion, and expansion in volume or cubical expansion. In the case of liquids and gases we are concerned only with their cubical expansion.

2. Change of temperature, of course, means simply that the effect of an addition of heat is to make the body get hotter and hotter, while a subtraction causes it to become colder and colder.

3. Change of state includes changes in the physical condition known as liquefaction or becoming liquid, and vaporisation or becoming converted into vapour. Thus, if we heat ice it first liquefies or becomes water, and is then vaporised or becomes steam.

CHANGE OF SIZE.

Expansion. EXPT. 125.—Take a metal ball suspended by a chain as shown in the figure, and suspend it by the side of a metal ring, through which it passes easily, or fits loosely. Heat the ball in a laboratory burner for a few minutes, and then try to drop it through the ring. It is too large and rests on the ring. Now allow it to cool slowly and notice that after a short time it gets smaller and will slip through quite easily.

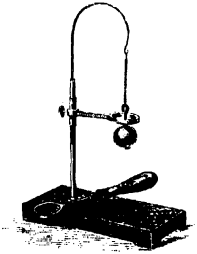
FIG. 53.—The Expansion of a Solid.

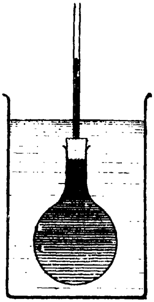
FIG. 54.—The expansion of a Liquid.

EXPT. 126.—Procure a 4 oz. flask and fit it with a cork. Bore a hole through the cork and pass through a long glass tube which fits tightly. *Fill* the flask with water coloured with red ink. Push the cork into the neck of the flask and so cause the coloured water to rise up the tube. See that there is no air between the cork and the water. Now dip the flask in warm water, and notice that the liquid gets larger and rises up the tube. Take the flask out of the warm water, and see that the coloured water gets smaller as it cools and that it sinks in the tube. See Fig. 54.

EXPT. 127.—Procure a well-made paper bag and tightly tie a piece of tape round the open end. Hold the bag in front of the fire and notice that the air inside gets larger and inflates the bag.

These experiments convince us that all bodies, whatever their physical condition, solid, liquid, or gaseous, get larger when heated and smaller when cooled. Now if they get larger their volume increases, and as the amount of matter in them, *i.e.*, their mass, remains the same, from what we have learnt about density it is clear that it must become less, or *bodies get lighter, bulk for bulk, when they are heated.* The converse holds true, *if they get cooler their density gets greater.*

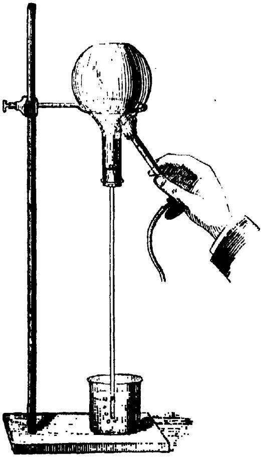

Fig. 55.—The Principle of the Air Thermometer.

Other Experiments to Illustrate the Expansion of Bodies when Heated.—EXPT. 128.—Solder together side by side a brass wire and an iron wire, each about two feet long. Hammer the compound wire straight, and notice how it bends when heated.

EXPT. 129.—Fuse a piece of platinum wire through the side of a glass tube, and notice that the glass does not crack on cooling. The platinum and glass expand about the same amount for a given increase of temperature.

EXPT. 130.—Fit with corks the necks of three 4 oz. flasks. Fit tightly into the corks three narrow glass tubes open at both ends. Fill the flasks with water, alcohol, and oil of turpentine respectively. Push in the corks till the liquid stands in each tube at the same height. Put all three flasks to the same depth into a vessel of warm water. Notice that the expansion of the glass causes a momentary sinking of the liquids ; that ultimately the expansions are very different.

EXPT. 131.—Fit a 2 oz. flask with a tight cork through which a tube passes, the upper end of which is bent down and then up at the end. Clamp the flask so that the end of the tube dips under water in a basin. Fill a test-tube with water, and invert it over the end of the tube. Warm the air in the flask, and collect the expelled air in the test-tube.

EXPT. 132.—Tightly fit a cork, through which a straight tube passes, into the neck of a 2 oz. flask. Turn over and pass the tube through the cork in the neck of a wide-mouthed bottle, containing coloured water. Warm the flask with the hand or a flame so as to expel some of the air, and let the liquid rise in the stem, Fig. 55. This constitutes an *Air Thermometer*.

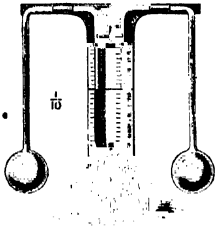

FIG. 56.—To illustrate Experiment 133.

EXPT. 133.—Fasten two bulbs or flasks together (air-tight) by a tube bent six times at right angles, and containing some coloured liquid in the middle bend, Fig. 56. Show that the liquid moves if one flask is warmed more than the other.

CHANGE OF TEMPERATURE.

Its Measurement.—The change of size which bodies experience when heated can evidently be made to provide us with a method of *measuring the change of temperature* which they undergo. If we can find some form of matter which expands regularly as it is heated, the increase in volume which results can be taken as a measure of the change of temperature. Thus, using the arrangement in Expt. 126, if we notice that the coloured water in the tube rises through a certain number of inches after being heated for some time, we can look upon this rise of the level of the water in the tube as an equivalent of a certain change of temperature, and we can be sure that if as the result of its contact with any other body, the liquid rises through this amount, that the body with which it is in contact has caused it to experience the change of temperature to which it is equivalent, and that the body is at the temperature represented by the higher level of the water. An arrangement of this kind is, then, a "temperature measurer," or, as it is always called, a **thermometer.**

It has been found by a long series of experiments that liquids expand much more than solids for a given change of temperature, and that gases again expand still more than liquids. Further, it has been observed that **all gases expand to the same amount for a given change of temperature.**

Since gases expand so much more than liquids, thermometers in which some gas is the form of matter which expands will be very much more accurate than liquid thermometers. Instruments are actually made in which air is used. They are called *Air Thermometers*, and are very accurate indeed.

Thermometers. *Choice of Materials.* (1) *Liquid.*—The property of substances which we have now considered is made use of in the construction of temperature measurers or thermometers in the following way. First we have to choose a suitable substance, and generally one of two liquids is selected, depending upon the particular purpose to which the thermometer is to be put. If it is to be employed for the measurement of very low temperatures, we shall use alcohol in our thermometer since it is a liquid which freezes with difficulty. If we wish to measure higher temperatures, alcohol would be unsuitable since it boils at a temperature of 78° C.[1] which is, as we shall see, considerably lower than the boiling point of water. For such higher temperatures mercury is used because it does not boil until the temperature of 357½° C. is reached. But though its boiling point is high enough to make it very valuable for the purpose named, it cannot be used for measuring very low temperatures, as at −40° C. it solidifies.

There are other reasons for selecting mercury in addition to the important one we have just given. It is a liquid whose level can be easily seen; it does not wet the vessel in which it is contained; it expands a considerable amount for a small increment of temperature; it is a good conductor of heat, and consequently it very quickly assumes the temperature of the body with which it is placed in contact. Very little heat is required to raise its temperature, and there is therefore very little loss of heat due to warming the thermometer.

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[1] These numbers will be understood after the section on the graduation of a thermometer has been read.

(2) *The Tube.*—Liquids, as we know, must be contained in vessels to keep them together. Consequently, the liquid we have chosen will have to be enclosed in some sort of case. For this purpose a glass tube is used, and some care must be exercised in selecting one. It should have a narrow bore, so that a given expansion of the liquid shall be spread over a considerable distance. The bore must be of equal size, or as nearly as we can get it, in every part, and as it is impossible to get one which is quite regular—it is usual for very accurate work to *calibrate* it, *i.e.*, to find the volume of each part of the bore.

Construction of a Thermometer.—Having selected a suitable piece of thermometer tubing, a bulb must be first blown on one end. The glass is melted at this end and allowed to run together and so close up the bore, and while the glass is still molten, air is blown down the tube from the other end, keeping the tube moved round, so that the bulb is symmetrically placed with reference to it. The bore of the tube is so fine that it is impossible to pour the liquid down it; some other plan must therefore be adopted. The tube is warmed and inverted in some of the liquid. Let us suppose we are using mercury. Warming the tube makes the air inside it expand, and of course some is driven out. As the tube cools the mercury is forced in by the weight of the atmosphere to fill the place of the expelled air. By repeating this alternate process of warming and cooling, under the circumstances we have described, enough mercury is soon introduced into the tube. The next step is to seal up the tube, leaving no air above the mercury; to do this the bulb is heated to a temperature slightly higher than we shall want our thermometer to register, the mercury

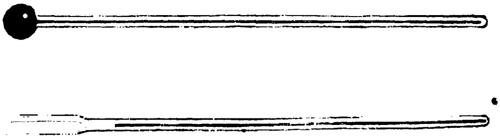

FIG. 57.—Thermometers before Graduation.

expands, and when it has reached the top of the tube, the end is closed by directing a blow-pipe flame against it. This method of closing a tube and keeping the air out is called *hermetically* sealing it. The thermometer at this stage should be put on one side for some days at least, in order that it may assume its final size, which it does very slowly indeed. The thermometer will now look like one of those in Fig. 57, according to the shape of the bulb which has been blown.

Graduating the Thermometer—Fixed Points.—In the graduation of a thermometer the plan always adopted is to choose "two fixed points" from which to number our degrees of temperature. The most convenient lower fixed point we can get is the temperature at which ice melts, or water freezes, for this is always the same if the ice is pure, and remains the same as long as there is any ice left unmelted. The truth of this statement can be verified with a thermometer in the condition shown in Fig. 57. Whenever it is put into melting ice the mercury always stands at the same level, or, as we have seen, melting ice is always at the same temperature and may be used to give one fixed point. The "higher fixed point" chosen is that at which pure water boils at the sea-level. We have to make this stipulation, for the boiling point of a liquid is altered when the pressure is changed, being raised if the pressure is greater and lowered if the pressure is less. When the water boils the temperature of the steam is the same as that of the water, and remains so as long as there is any water left. The lower fixed temperature we refer to as the "Freezing Point of Water," the higher as the "Boiling Point."

Marking the Freezing Point.—For this purpose an arrangement like that shown in Fig. 58 is very suitable. The funnel is filled with pounded ice, which before powdering had been carefully washed ; or snow might, if more convenient, be used. The glass dish catches the water which is formed from the melting of the ice or snow. We make a hole in the pounded ice by thrusting in a pencil or glass tube about the size of the thermometer, and into this hole we put the thermometer and support it so that the whole of the mercury is surrounded by the ice or snow. The arrangement is left for about ten or fifteen minutes, until it is quite certain that the tube and mercury are at the same temperature as the melting ice. When this is so the tube is raised until the mercury is just above the ice, and a fine scratch made with a three-cornered file on the tube at the level of the mercury.

Marking the Boiling Point.—The apparatus used for doing this will be readily understood by examining Fig. 59. Water is contained in the cylinder E, which is of a much smaller diameter towards the top, where it is open in the way the drawing shows. Over this smaller open cylinder is placed a larger one, which rests on the larger cylinder containing the water, and is provided with three openings—one at the top A through which to pass the thermometer, one at the side C to allow superfluous steam to escape, and one for the introduction of a bent tube containing mercury, for indicating whether the pressure inside the apparatus is the same as that of the surrounding air, and which is known as a *manometer*.

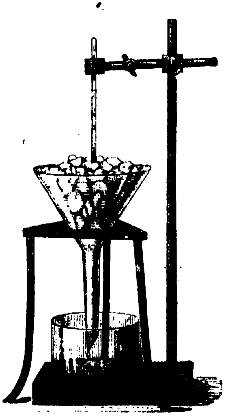

FIG. 58.—Marking the Freezing Point.

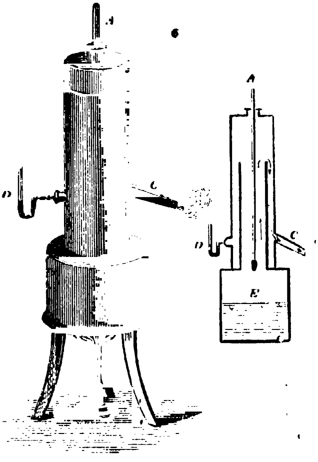

FIG. 59.—Marking the Boiling Point.

The apparatus is arranged as shown in the left figure (Fig. 59), and the water in E boiled. Steam is generated, and circulates as shown by the arrows. The thermometer is exposed to the steam until it has taken the temperature of the boiling water, and the level of the mercury is then scratched on the tube.

Choosing the Scale.—Some value must now be given to these two fixed points, and of course they can be called anything the maker likes, but for the sake of comparing one man's observations and experiments with those of other people it is most convenient to graduate all thermometers in the same way. The thermometers of this country are divided up in two ways—(1) the Centigrade scale, (2) the Fahrenheit scale.

The Centigrade Scale.—Here the freezing point is called *zero* or *no degrees*, written 0° C. The boiling point is called *one hundred degrees Centigrade*, and is written 100° C. The space between these two limits is divided into 100 parts, and each division called a *degree Centigrade.*

The Fahrenheit Scale.—On thermometers marked in this way the freezing point is called *thirty-two degrees Fahrenheit*, written 32° F., and the boiling point *two hundred and twelve degrees Fahrenheit*, written 212° F. The space between the two limits is divided into 180 parts and each division is called a *degree Fahrenheit.* The reason of this difference is interesting. The physicist Fahrenheit, after whom the thermometer is named, got, as he thought, a very low temperature, by mixing common salt with the pounded ice when measuring the lower fixed point, and he imagined that he had got the lowest temperature which could be reached, and called it zero. His conclusion was wrong, and the mistake has brought about two ways of measuring temperatures.

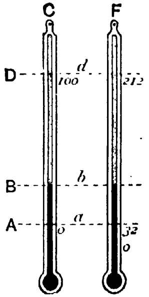

FIG. 60.—The Fahrenheit and Centigrade Scales.

Conversion of Scales.—It should be clear from what we have said that the interval between the boiling and freezing points, that is, the same temperature difference, is divided into 100 parts on the Centigrade scale and 180 parts on the Fahrenheit, and consequently 100 Centigrade degrees are equal to 180 Fahrenheit degrees, which is the same as saying one degree Centigrade is equal to nine-fifths of a Fahrenheit degree, or one degree Fahrenheit is equal to five-ninths of a degree Centigrade.

$$100 \text{ C degs.} = 180 \text{ F degs.}; \therefore 5 \text{ C.} = 9 \text{ F.} \therefore \text{C.} = \tfrac{9}{5} \text{ F. or F.} = \tfrac{5}{9} \text{ C.}$$

In converting Fahrenheit readings into Centigrade degrees, we must subtract 32 (because of what has been said of the freezing point on the former scale) and multiply the number thus obtained by 5 and divide by 9. To change from Centigrade to Fahrenheit. multiply the former reading by 9 and divide by 5 and add 32 to the result.

> EXAMPLE.—What temperature on the Fahrenheit scale corresponds to 20° C?
>
> ANSWER.—20° C is 20 C degs. above temperature of melting ice, *i.e.* $20 \times \frac{9}{5}$ Fahr. degs. above 32° F = $(36+32)^\circ$ F = 68° F.

When it is necessary to refer to temperatures lower than the freezing point of water a minus sign is placed before the temperature, thus three degrees below the freezing point of water on the Centigrade scale is written -3° C.

Distinction between Heat and Temperature.—Temperature is not heat. It is only a state of a body, for the body may be cold one minute and hot the next. A hot body is one at a high temperature, a cold body one at a low temperature. If a hot body and a cold body be brought into contact there is a passage of heat from the hot one to the cold until they are both of the same degree of hotness or coldness. Now in this last sentence substitute "at a high temperature" for "hot," and "at a low temperature" for "cold," and we shall see our way to a definition of temperature; thus, if a body at a high temperature and one at a low temperature be brought into contact there is a passage of heat from the former to the latter until they are both at the same temperature. Hence, we can define temperature as **a condition of bodies that determines which of two bodies when placed in contact will part with heat to the other. Temperature may also be defined as intensity of heat.**

Evidently temperature is analogous to the level of water, for we have learnt that if two cisterns containing water at different levels be put in connection there will be a flow of water from the one where the water stands at the higher level to the other until they assume the same level.

This analogy between temperature and water-level is worth considering more fully, as much can be learnt about the distinction to which the heading refers. Imagine a cylinder of water; the most accurate determination of the height at which the water stands gives no information concerning the *quantity of the liquid* contained by the cylinder; similarly, a complete knowledge of the temperature of a body tells nothing about the *quantity of heat* which can be got out of it. Just as the water-level and temperature are comparable terms, so the quantity of water in a vessel can be compared with the quantity of heat in a body. To ascertain the quantity of water in the cylinder, in addition to knowing the height of it, the capacity of the cylinder for water or its size must be known: so also to measure the quantity of heat in a body, in addition to knowing its temperature, we must also be informed of its **capacity for heat.**

EXPT. 134.—Arrange three glass cylinders, of different diameters but equal heights, in a row. Pour a wine-glassful of water into each of them in succession. Notice that the same quantity of water fills the cylinders to different heights. The level is highest in the cylinder of smallest diameter and lowest in the one of largest diameter. The capacity for water of the one with the greatest diameter is evidently more than in either of the other cases, and consequently the same quantity of water fills them to extents inversely proportional to their capacities for water.

EXPT. 135.—Mix 1 lb. of hot water with 1 lb. of cold, and observe the temperature of the mixture. This temperature will be found half-way between the two original temperatures.

EXPT. 136.—Heat equal weights of lead and water in the same beaker. Provide two other beakers containing equal weights of cold water. Put the hot lead in one of these, the hot water into the other. Stir and note the temperatures. Observe carefully that the water into which the heated lead is plunged is not at so high a temperature as in the other case. Equal amounts of water at the same temperature are thus shown to be heated to different extents by equal weights of water and lead at the same high temperature.

EXPT. 137.—Mix 1 lb. of mercury at 100° C with 1 lb. at the temperature of the room. Notice that the resulting temperature is midway between the temperatures of the two quantities of mercury.

EXPT. 138.—Mix 1 lb. of water at the temperature of the air with 1 lb. of iron at 100° C, and notice the temperature of the mixture. Mix 1 lb. of water at 100° C with 1 lb. of iron at the atmospheric temperature, and notice that the resulting temperature is in the latter case much the higher.

EXPT. 139.—Shake up known weights of water and mercury at different temperatures, and note the resulting temperatures. Repeat, using water and turpentine at known but different temperatures. Confirm by shaking up mercury and turpentine. The relative capacities for heat are inversely as the weights, and inversely as the change of temperature.

Measurement of Quantities of Heat.—As in all other cases of measurement, we must have some unit in terms of which to compute the quantity of heat which is being measured. The **unit of heat** generally adopted **is the quantity of heat necessary to raise the temperature of 1 gram of water through 1° C.** In terms of this unit the amount necessary to raise 10 grams of water through 1° would be 10 units of heat, or that required to raise 1 gram through 10° would be the same amount. Water has the greatest capacity for heat of all forms of matter; in other words, the quantity of heat required to raise 1 gram of water through 1° C. is greater than the quantity required to raise 1 gram of any other substance through the same interval of temperature.

As we have defined the unit quantity of heat, the capacity which any substance possesses for receiving heat may be expressed in terms of the unit. **The capacity for heat of a body is the number of units of heat required to heat it through 1° C. The capacity for heat (or thermal capacity) of unit mass is called the Specific Heat.**

As water has the highest capacity for heat, any weight of water at any temperature contains more heat than the same weight of any other substance at the same temperature. A pound of water at 100° C would thus possess a greater heating effect than a pound of lead, or iron, or copper at the same temperature. It has been seen by experiment that when equal weights of the same substance are mixed together, the temperature of the mixture is half-way between the temperatures of the two parts which make it. If, however, equal weights of two different substances at different temperatures are mixed together, the resulting temperature lies nearer the temperature of the substance with the greater specific heat. In estimating the quantity of heat taken up or given out by a body, we must find the product of three quantities, viz., its mass, its capacity for heat, and the range of temperature through which it is heated or cooled respectively.

In considering any question on capacity for heat, it should be borne in mind that the amount of heat given out depends upon the weight of the substance involved, its change of temperature, and the heat capacity of unit mass, or its specific heat.

Conduction of Heat.—By touching a succession of things in a room, say the marble mantel-piece, the fender, the back of a chair, the hearth-rug, we obtain a succession of sensations: the first two we say are cold, the chair-back not quite so cold, while the rug feels quite warm, and yet they are one and all under the same conditions and there is no reason why they should not be at the same temperature. The explanation of these different sensations is really very simple. In all those cases where the hand *receives heat* we feel the sensation of warmth, while in those where the hand *gives out* heat we say the body is cold or cool. Fig. 61 will enable the student to remember this. Now we see why the fender feels colder than the hearth-rug. The fender takes more heat from the hand than the hearth-rug, and it does so because it is a better *conductor of heat*.

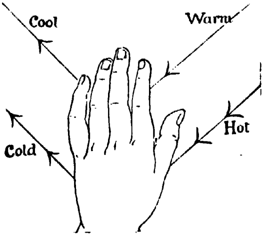

FIG. 61.—Hot and Cold Bodies.

We shall do well to consider this expression a little. Put one end of a poker in the fire and hold the other. Soon the poker begins to feel warm, and as time goes on it gets warmer and warmer, until at last you can hold it no longer. Heat has passed from the fire along the poker, or has been *conducted* from the fire by the poker.

The process by which heat passes from one particle of a body to the next is called conduction, and the body along which it passes is known as a conductor.

Good and Bad Conductors of Heat.—As we have learnt, heat is conveyed by *conduction* from one particle of a body to the next, the heat travelling from the hotter to the colder parts.

Such conduction causes no visible movement of the parts of the body. This is the process by which heat passes through solids. Those substances which easily transmit heat in this way are called **good conductors,** while those which offer a considerable amount of resistance to its passage are called **bad conductors.**

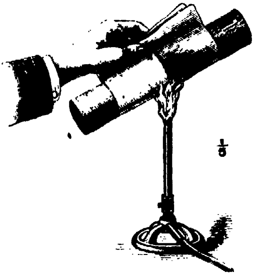
Fig. 62.—Experiment 140.

EXPT. 140.—Wrap a piece of paper smoothly round a brass tube and hold in the flame of a gas burner. The paper is not scorched. Wrap the paper around a wooden rod of the same size, and heat as before: the paper is scorched, Fig. 62. Brass is a good conductor, wood but a poor one. How does this explain what you have noticed?

EXPT. 141.—Twist an iron and a copper wire together at one end. At about four inches from the joint fasten a marble on each with beeswax. Heat the joint in a Bunsen flame. Notice that the marble on the copper is the first to fall. Why?

EXPT. 142.—Make a short coil of stout copper wire ¼ inch internal diameter. Pass it over the wick of a candle without touching the wick. The candle is extinguished owing to the cooling effect of the wire which conducts away the heat.

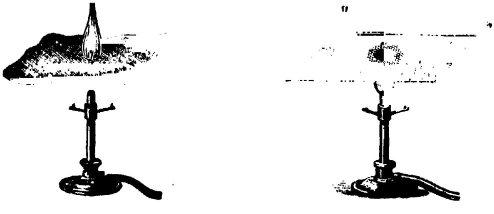
FIG. 63.—Experiment 143.

EXPT. 143.—Turn on, but do not light, a gas jet. Hold over it a wire gauze, and light the gas above the gauze. Notice that the flame does not strike through. Why? Vary the experiment by lowering a piece of cold wire gauze upon an ordinary Bunsen flame. What happens?

Liquids are bad conductors of heat. We can show this in the case of water by the following experiment.

Expt. 144.—Fill a test-tube three quarters full with cold water, and having weighted a small piece of ice by winding wire round it, or in some other way, drop it into the test-tube. Hold the test-tube near the bottom where the piece of ice is, and warm the top of the water in a Bunsen flame, as shown in Fig. 64. The water at the top can be heated until it boils vigorously and yet the ice is not melted, showing what a bad conductor the water is.

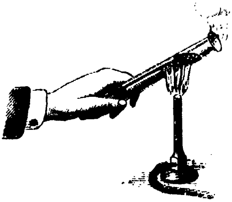

Fig. 64.—To illustrate that Water is a bad Conductor of Heat.

Liquids are bad conductors of heat, but gases are far worse, as the following experiments indicate :—

Expt. 145.—Examine the shadow of a red-hot poker. Notice that the heating of the air as exhibited by its flickering extends but a very little way downwards, thus showing that air is a bad conductor of heat.

Expt. 146.—Place a little lime in the palm of the hand and bring the point of the hot poker upon it. The air enclosed in the lime does not conduct the heat of the poker, so the hand is not burnt.

Everyday applications of these facts about Bad Conductors.—To keep ice in the warm days of summer the custom is to wrap it up in flannel and put it into a refrigerator. The flannel, because of its loose texture, encloses a quantity of air, which, being a bad conductor of heat, prevents the passage of heat from the warm outside air to the cold ice inside.

Similarly, ice which has to be conveyed by rail or boat is packed in sawdust.

The refrigerator itself, too, depends upon much the same facts. The common form consists of a double-walled box with a space between the walls. This is either left "empty" as it is called, when it is full of air; or, it is filled with some other bad conductor, such as the mineral substance *asbestos*.

If we wish to lift a hot plate we hold it with a folded cloth for the same reason. Cylinders of engines are encased in a packing of some badly conducting material.

Convection. How Liquids and Gases become heated.—The process by which water and other liquids are heated must now be considered. The water nearest the source of heat gets heated, expands, and in consequence gets lighter; it therefore rises through the general mass of the liquid. Something must take its place, and the cold water at the top being heavier sinks and occupies the space of the water which has risen. This water in its turn gets heated and rises, and more cold water from the surface sinks. This gives rise to upward currents of heated water and downward currents of cool water, until by and by the whole of the water is heated. These currents are known as **convection currents**, and the process of heating in this manner is called *convection*.

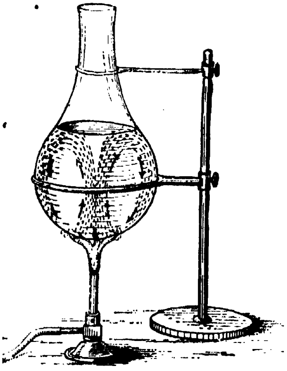

FIG. 65.—Convection Currents.

EXPT. 147.—Heat over a small flame a round-bottomed flask full of water, as in Fig. 65. Throw into the water some solid colouring matter, like cochineal, aniline dye, litmus, &c. Notice how the hot and coloured water ascends.

Gases are similarly heated by this process of convection, which may be thus defined:—**Convection is the process by which fluids (liquids and gases) become heated by the actual movement of the particles of the fluid.**

Applications of Heating by Convection. Heating Buildings by Hot Water.—One of the commonest ways of heating large buildings is by means of hot-water pipes, and the efficacy of this plan is due to the facts we have just learnt. In Fig. 66 we have the condition of things in such a building very simply represented. We will suppose, to begin with, that the boiler B, the pipes *ab*, &c., and the coils C, C′, are nearly full of cold water, and that the fire below the boiler is lighted. Heat passes through the bottom of the boiler by the process of conduction, and heats the layer of water near it, which, expanding, rises and passes up the tube *ab* to the top of the building, where it gives out its heat to the rooms. The place of this water which has thus risen is taken by cold water from the other pipe terminating just past *d*. This in its turn gets warmed and rises, and its place is taken by the water which has become cold by its passage through the pipes in the various rooms. There is thus no difference in this case from what we have seen to be true in the flask of water being heated from below.

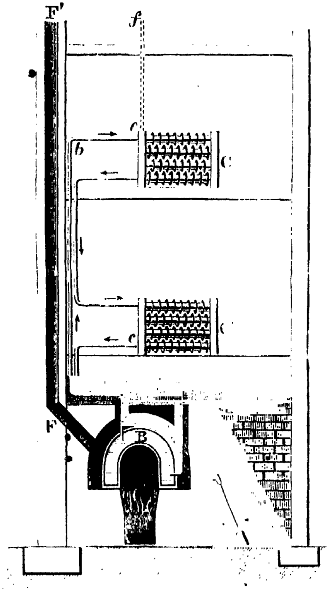

FIG. 66.—Heating a large building by hot water pipes.

Ventilation.—The ventilation of ordinary dwelling rooms is easily possible because of the way in which gases become heated. The air in a room becomes warmed and rendered impure at the same time. Consequently there is a tendency for the vitiated air to rise, and if a suitable place near the ceiling is made for its egress as well as a place near the floor for the colder, purer, air from outside to enter, we shall have a continuous circulation of air set up which will keep the atmosphere of the room pure and sweet.

Expt. 148.—Slightly open the door of your room, and hold a lighted candle (*a*) near the bottom, (*b*) at the middle of the opening, (*c*) near the top. How is the light affected? (Fig. 67).

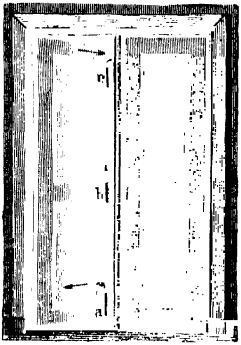

Fig. 67.—To illustrate Experiment 148.

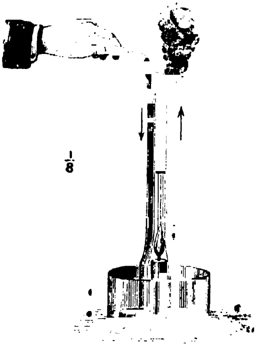

Fig. 68.—How to ventilate.

Expt. 149.—Place a short piece of candle in a saucer, light it, put a lamp glass over it, and pour sufficient water into the saucer to cover the bottom of the lamp glass (Fig. 68). Watch how the light of the candle is affected and describe what happens. Next cut a strip of card less than half the height of the lamp glass, and nearly as wide as the internal diameter of the top. Insert the card into the lamp glass so as to divide the upper part into two halves. Now light the candle again, and see whether it will burn with the divided chimney over it.

Test the direction of the currents of air at the top of the chimney by holding a smoking taper or match over it.

CHANGE OF STATE.

Melting Point.—The first effects on heating a solid, as has been seen, are to raise its temperature and to cause an increase in its volume. If we continue the process of heating, the solid state gives place to a liquid condition, or the solid melts or fuses. Common experience affords many examples of this, *e.g.*, when a lump of lead is heated its temperature rises and it gets larger, and as the heating is continued it is converted into a silvery liquid. Wax, ice, and iron are other examples. But the temperature at which the liquid state is assumed, *i.e.*, the **melting point,** is widely different in the case of different substances, as the following table shows :—

Examples of Melting Points.

Ice	0° C.	Tin	230° C.
Beeswax	65°	Lead	330°
Sulphur	115°	Cast Iron ...	1200°

EXPT. 150.—Pound some pieces of clear ice and thrust a thermometer into the powdered mass. Record the temperature indicated by the thermometer.

Put some of the ice into a beaker and pour in some water, stir the mixture and again record the temperature.

Place the beaker on a piece of wire gauze or in a sand bath and warm gently. Notice the reading of the thermometer *as long as there is any ice unmelted.* In all these cases the temperature is the same, or the temperature of melting ice is constant.

EXPT. 151.—Soften a piece of glass tubing in the flame of a blowpipe and, removing it from the flame, draw it out until the bore becomes very small. Cut off a piece of it and dip it into some melted beeswax. In this way the fine tube becomes filled with wax which soon solidifies. Tie this filled tube on to a thermometer, near its bulb, and put the thermometer into a beaker of water which has been placed over a burner in the way shown later in the Chapter on Sulphur (Chapter XVII.) Gradually heat the water and notice when the wax melts, and at that instant read the thermometer. This reading will be the *melting point* of the wax.

Boiling Point.—In explaining the nature of the process by which liquids are heated, it was seen that the bottom layers of water become heated and rise, their place being taken by the colder layers from above. After this heating has been continued for a certain length of time, the bulk of the water gets so hot that the bubbles of vapour which are formed near the source of heat are not condensed again in their upward passage through the liquid, and coming to the surface they escape as steam. The temperature at which bubbles of this sort get formed throughout the mass of the liquid is quite definite for a given pressure of the atmosphere and is known as the **boiling point.**

EXPT. 152.—Fit a cork, with two holes through it, into a flask or test-tube (Fig. 69). Half fill the flask or test-tube with water, and push the thermometer through one of the holes in the cork until the bulb is wholly immersed. Into the other hole fit a piece of glass tubing bent at right angles. Place the flask or test-tube upon a retort stand, and observe the division of the thermometer scale level with the top of the mercury. Gently heat the water until it boils, noticing how the mercury of the thermometer is affected throughout the process. See where the mercury stands when the water is boiling and find whether any difference is produced when the water is boiling furiously and when it is boiling gently. Raise the thermometer until the bulb is just above the boiling water, and observe the temperature which it then indicates.

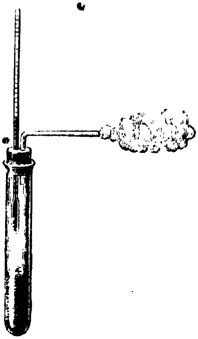

FIG. 69.—Determining a boiling point.

Effect of Pressure on the Boiling Point.—A word or two must be said with respect to the reservation which has been made about the pressure of the air. It has been seen that the weight of the atmosphere is very considerable. It presses upon the surfaces of all bodies with a force dependent upon its weight, which, like all other forms of matter, is proportional to its mass, and the mass of the atmosphere will depend upon the extent of the air above the body, which will clearly be less at the top of a mountain than at the bottom of a mine.

If we wish to boil a liquid, therefore, in those cases where the pressure of the atmosphere is great we shall have to heat the liquid more before the bubbles of vapour can escape at the surface than when the pressure is less. If we heat the liquid more its temperature will get higher before there is any conversion into vapour, and consequently its boiling point will be higher when the pressure is greater. In finding the boiling point of a liquid we must therefore know the pressure of the atmosphere at that place and time.

> EXPT. 153.—Get a well-fitting, sound cork to fit the neck of a 4 oz. round-bottomed flask. Boil water in the flask till all the air is expelled. Remove the burner and cork up the flask as rapidly as possible. Turn the flask over and pour cold water on the up-turned bottom. Notice that as the pressure is diminished by the condensation of the vapour the water boils again at a lower temperature.

Latent Heat of Fusion.—It has been already pointed out that when ice is melting the mixture of ice and water remains at the same temperature as long as any ice is left, although the mixture is being continuously heated. The heat which is thus absorbed, and has no effect on the temperature of the mixture, is used up in causing the change of state from solid to liquid, and since it lies hidden, as it were, is referred to as the *latent heat of fusion* or the *latent heat of water*. This quantity of heat can be defined as being **the number of units of heat which are required to change the state of the unit mass of ice, converting it from the solid to the liquid condition, without raising its temperature.**

To melt one pound of ice requires as many heat units as are necessary to raise a pound of water from 0° to 80° C., or as much heat as is wanted to raise 80 lbs. of water through 1° C. In just the same manner, before a pound of water can be changed into a pound of ice we must take away from it just the same amount of heat. Bearing this in mind, it is easy to understand why it takes so many cold nights to cover a pond over with ice, and why it takes such a long time, too, to completely melt the snow in our roads even after the thaw has set in.

Latent Heat of Vaporisation.—When a liquid boils, as we have seen, the thermometer remains stationary as long as there is any liquid left. After boiling has commenced all the heat is absorbed in changing the substance from the liquid to the gaseous state. This heat also becomes "latent" and is refered to as the *latent heat of vaporisation* or the *latent heat of steam*; it can be defined as **the number of heat units required to change the unit mass of water into steam without changing its temperature.**

The latent heat of water can be measured by mixing a known quantity of ice with a known quantity of warm water and ascertaining the amount of heat given up by the warm water to melt the ice. In a similar manner, by mixing a known quantity of steam with a known amount of water and ascertaining how much heat the steam gives to the water, we can measure the latent heat of steam.

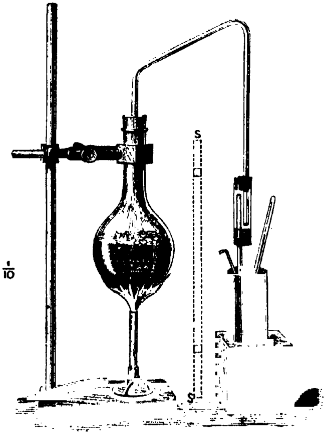

FIG. 70.—Arrangement for passing steam into a vessel containing cold water.

EXPT. 154.—Put some ice into a known weight of hot water. Notice the temperature when the ice is melted. Determine the weight of the ice by reweighing the water. Notice that a small quantity of ice requires a large quantity of heat to melt it.

EXPT. 155.—Pass steam, produced by heating water in a flask, through a delivery tube into cold water. Notice the rapid rise in temperature of the water, and determine the weight of water condensed. To prevent hot water from coming out of the end of the delivery tube, a trap such as is shown in Fig. 70 should be used.

There is an absorption of heat which has no effect on the thermometer whenever a liquid is converted into a vapour.

We have taken the case of water as an example, but the same general phenomenon can be noticed in every case.

EXPT. 156.—Pour a few drops of any very volatile liquid, such as ether or carbon bisulphide, upon the hand. It soon disappears, and as a result of its disappearance the hand feels very cold. The heat necessary to effect the change of state has been taken from the hand.

EXPT. 157.—Place a few drops of water between the bottom of a small beaker and a block of wood. Put some ether into the beaker and blow over it with a pair of bellows having a tube fastened to the nozzle. The water will freeze.

When the weather is hot it is a common plan to wrap wet cloths round bottles of wine to keep the beverage cool. In this case the evaporation of the water on the cloth is possible only by the absorption of a large amount of heat, which is extracted from the bottle. For precisely the same reason, water is kept in porous vessels in hot countries.

CHIEF POINTS OF CHAPTER X.

Effects of Heat.

Change of Size.	*Change of Temperature.*	*Change of State.*
Substances expand when heated and contract when cooled, the amount of change varying with different materials.	When a body is gaining heat it rises in temperature, and when it is losing heat it falls in temperature.	By the addition of heat a solid can be melted into a liquid, and then vaporised; the changes happen in the reverse order if sufficient heat is withdrawn.

A Thermometer is an instrument for measuring temperature. Its action usually depends upon the fact that substances expand when heated and contract when cooled.

Why Mercury is generally used in Thermometers.—The following are the chief reasons why mercury is the best liquid to use in thermometers under ordinary circumstances :—

(1) It remains liquid through a long range of temperature (−40° C. to 357° C.).

(2) It expands considerably for a small rise of temperature.

(3) Its amount of expansion for one degree rise of temperature is fairly uniform while it remains liquid.

(4) It is a good conductor of heat, and therefore it quickly assumes the temperature of objects surrounding or touching it.

(5) Comparatively little heat is required to raise its temperature, that is, it has a low specific heat.

(6) It can be easily seen, and does not wet the glass containing it.

Thermometric Scales.—The distance between the marks of freezing and boiling points on the stem of a thermometer may be divided as follows:—

	Fahrenheit scale.	Centigrade scale.
Boiling point	212°	100°
Freezing point	32°	0°

Discrimination between Heat and Temperature.

Heat.	*Temperature.*
Dependent on the energy of the vibration of the particles of a body.	Dependent on the velocity of the vibration of the particles of a body.
Capable of producing change of temperature.	Is changed by the addition or withdrawal of heat.
Measured in heat-units.	Measured by degrees.
Analogous to water flowing from high to lower levels.	Analogous to difference of level.

Capacity for Heat.—Some substances have a greater capacity for heat than others. This is analogous to vessels of different sectional area.

Capacity to hold Liquids.	*Capacity for Heat.*
The same quantity of water poured into a narrow and into a wide jar produces different changes of level.	The same quantity of heat produces different temperatures (heat level) in equal masses of different substances.
A wide jar contains a greater quantity of water than a narrow one filled to the same height.	A body having a large heat capacity, if at a given temperature, contains a greater quantity of heat than the same weight of a substance of less capacity for heat at the same temperature.

Quantity of Heat.—The amounts of heat absorbed (or given out) by equal weights of different materials when heated (or cooled) through the same range of temperature are in general different.

The Capacity for Heat or the Specific Heat of a body is the number of units of heat required to heat one gram of a substance through 1° C.

Transmission of Heat.—Heat travels from one body to another by conduction, convection, or radiation.

Conduction.	*Convection.*	*Radiation.*
The particles pass on the heat vibrations from one to the other. This is how solids get hot.	Hot particles move away from the source of heat, and cooler ones take their place. This is how liquids and gases chiefly get heated.	Due to the vibrations of a heated body being transmitted through the ether: example, the heat of the sun transmitted through space to the earth.

Change of State.—Many solids change into liquids, and liquids into gases, at definite temperatures. During these changes of state heat is absorbed without rise of temperature. The change of state generally involves a change of volume.

Solids + latent heat of liquefaction = Liquids.

Liquids + latent heat of vaporisation = Gases.

QUESTIONS ON CHAPTER X.

1. (*a*) Why is the tube of a thermometer made with a narrow bore, while the bulb is large?
   (*b*) Why is the top of a thermometer sealed up?
   (*c*) Why is mercury the best liquid to use in a thermometer under ordinary circumstances?
   (*d*) Under what circumstances is alcohol used instead of mercury?
2. Enumerate the effects which ensue when a solid body is continuously heated. How would you show these effects experimentally?
3. Describe experiments to prove the expansion by heat of (*a*) a solid, (*b*) a liquid.
4. How are the "fixed points" of a thermometer determined?
5. Draw some distinctions between heat and temperature.
6. Describe an experiment to show that different substances at the same temperature have different quantities of heat.
7. State the meaning of "conduction" and "convection" as applied to heat.
8. How would you prove that when ice is melted heat is absorbed without rise of temperature?
9. State precisely what is meant by the statement that the specific heat of mercury is ·033.

10. Describe an experiment to show that water is a bad conductor of heat.

11. Explain fully the manner in which liquids become heated. Describe some useful application of the facts.

12. Describe an effective plan for ventilating a dwelling room and point out the principle upon which it depends.

13. Define what is understood by the melting point of a solid. Explain how you would proceed to determine this temperature in the case of beeswax.

14. What happens precisely when a liquid boils? What connection is there between the boiling point of a liquid and the pressure of the atmosphere?

15. Give a detailed account of the changes in volume and temperature which are noticed as a piece of ice is continuously heated.

16. What is meant by the latent heat of fusion? Apply your definition to account for the length of time necessary to effect a complete thaw after a season of frost.

17. Describe and explain the means adopted to obtain ice in tropical countries.

18. Why is ice wrapped in flannel in the summer when we want to preserve it?

19. If a pound of water at 100° C. is mixed with a pound of water at 0° C°., the temperature of the mixture is 50°. How would the result have differed if a pound of oil at 100° C. had been substituted for the hot water? Explain the difference.

20. Four ounces of hot lead filings and four ounces of water at the same temperature are poured upon separate slabs of ice. Will the lead or the water melt the most ice? Give reasons for your answer.

21. What will be the temperature of the mixture produced by mixing:—(*a*) 1 lb. of water at 0° C. with 2 lbs. of water at 90° C.? (*b*) 1 lb. of mercury at 10° C. with 1 lb. of mercury at 100° C.?

22. Water sometimes spurts from the spout of a kettle standing upon a fire. How do you account for this, and how would you prevent it without taking the kettle off the fire?

23. What is a thermometer, and what information concerning heat does it supply?

Give an instance of each of the following effects produced by heat: (*a*) change of size, (*b*) change of temperature, (*c*) change of state.

24. Describe a differential air thermometer. How would you use it to investigate whether equal volumes of two different substances give out equal amounts of heat when cooled through the same range of temperature?

25. Why is a vessel of water heated more quickly if heat is applied at the bottom than if it is heated at the top?

Draw a diagram to illustrate the movements of a liquid heated from below.

26. Explain exactly how the boiling point of a thermometer is determined.

27. Point out the difference between the conduction and convection of heat.

Describe an experiment showing that water is a bad conductor of heat.

28. Describe carefully the manner in which the freezing and boiling points on a thermometer are determined.

29. Explain what is meant by specific heat. How would you show that equal weights of two different substances give out different amounts of heat when cooled through the same range of temperature?

CHAPTER XI

LIGHT

In Chapter IX the student has had his attention directed to this subject in considering energy. Radiation was there regarded as a form of energy, and was seen to present itself under several different aspects. Sound was instanced as an example of radiation which, resulting in the first place from a sounding body, is transmitted through the surrounding medium, air, as a vibration of its particles. The vibrating source of sound sets the neighbouring air particles oscillating backwards and forwards, and this one, in its turn, sets its neighbour in motion in the same way all along the line of travel of the sound disturbance. But after the passage of the disturbance the air returns to its previous condition and undergoes no permanent change. Such a mode of propagation is spoken of as *wave-motion.*

In this chapter, however, we shall chiefly concern ourselves with those radiations which include light and are conveyed through the other medium which has been already referred to. It is known as *the ether* and pervades all space; it exists as well throughout the mass of all material bodies. These radiations are also propagated by waves, though in a different manner from that in which sound travels in the air. They may be regarded as **ether-waves.**

Waves.—Before the nature of wave-movements in the ether can be understood, it will be necessary to learn what is meant by a wave, as well as those terms which are used in speaking about them, and this cannot be better done than by beginning with waves in water. Everybody has started these by dropping a stone into a still pond. Where the stone is dropped the water is pushed down into a hollow cavity which, as it is watched, spreads in every direction until the edge of the pond is reached. Had a cork been placed upon the water, it would have been noticed that when the disturbance reached it, all that happens to it is an up and down motion. It does not move forward with the wave. The downward motion of the water producing the depression when the stone comes into contact with the pond is followed by the production of a crest by the swinging back of the water particles, and this crest moves across the pond immediately after the depression and with the same velocity. This is succeeded by the formation of another depression, which is followed by a second crest, and so on until the effect of the impact of the stone has died away. The distance from one crest to the next or from one depression to the next is called a **wave-length**.

It must be very carefully noticed that the water particles themselves move *up and down*,[1] or vertically, as shown by the cork, whereas the wave itself moves along the surface of the water, or *horizontally*. These two directions are at right angles to one another, and such a wave is called a **transverse wave.**

If the motion of the particles is in the same direction as that of the propagation of the wave motion the wave is called **longitudinal.**

Ether-waves are of the first kind, *i.e.*, transverse waves; sound waves in the air are of the latter kind.

Ether-Waves are of Three Kinds.—The ether-waves which come to our planet from the sun, and comprise what is inclusively called sunlight, produce different effects. If they fall upon our bodies they may be absorbed, and the energy of the wave-motion become converted into heat, and for this reason the waves which have been absorbed are spoken of as **Radiant Heat**; if they fall upon the retina of an eye, they may produce a sensation of light, and we then call the waves **Light**; falling upon a photographic plate or upon a green leaf, they may produce certain chemical effects, and are then referred to as **Actinic rays.** The word "rays" means simply kinds of radiation, and hence we may speak of heat rays, light rays, and actinic rays. It cannot be too strongly insisted that in their passage through the ether these ether-waves do not give rise to any of these results; they are simply waves transferring energy by wave-motion.

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[1] Really each particle of water moves round in a circle, but the statement in the text brings out the nature of transverse waves.

SOME LAWS RELATING TO LIGHT.

Light Travels in Straight Lines.—That light rays travel in straight lines can be at once shown by examining the path of these waves as they pass through a hole in a shutter of a darkened room. Though, as has been learnt, the light-waves are not themselves visible, yet the path of the light becomes apparent because the minute particles of dust in the air are rendered luminous by the vibrations of the ether being absorbed by them. The path of the light can thus be followed, and is seen to be in a straight line. We can infer this from several everyday experiences. We cannot see round a corner ; if light travelled in lines that were sometimes bent (we are speaking of a uniform medium), there is no reason why we should not. Or, again, every one knows that it is only necessary to put a small obstacle in the path of the light from a luminous body to completely shut out our view of it. The light from the setting sun, when the sky is cloudy, is often seen to travel in straight lines.

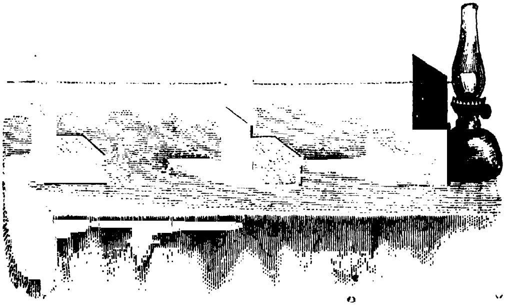


FIG. 71.—To illustrate that Light travels in Straight Lines.

EXPT. 158.—Take three cards and make a small hole in each with a fine needle. Fix the cards on to wooden blocks so that all the holes are at the same height and in a straight line. Place a lighted candle or a lamp in front of the card, and look through the third (Fig. 71). As long as the holes are in a straight line you can see the light from the candle shining through. Move one of the cards aside, and notice that you can no longer see the light. It must be remembered that what we have said about light applies equally to all other kinds of radiation.

EXPT. 159.—Place a stick vertically between the wall and a long thin luminous gas flame. Notice that the shadow of the stick is sharply defined.

EXPT. 160.—Construct a pin-hole camera as follows:—Make two paste-board tubes, by rolling pasted paper on a wooden cylinder, so that one fits inside the other. For the wider tube previously cover the cylinder with dry paper. Cover one end of the narrower tube with tissue paper and thrust this end into the wider tube. Line with black paper. Notice pictures on the tissue paper. Reason out how they are formed. Many toy-shops sell cheap pin-hole cameras, which may be used instead of constructing one from tubes.

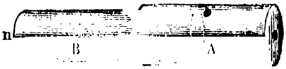


FIG. 72.—A Pin-hole Camera.

EXPT. 161.—Cover a lantern cap with tinfoil, remove the condensing lens, and place the cap on. Make pin-holes in the cap. For every pin-hole there is an image formed on the screen. Make the pin-holes more and more numerous, and near together, till the images overlap and become confused. At last diffused light is produced, which is an overlapping of images.

Reflection of Light.—When any wave is said to be *reflected*, it is understood that it comes into contact with the surface of some body, and is thrown back from that surface, and travels in a direction opposed to that in which it was originally moving. This may happen in two ways, either **regularly** or **irregularly**. In the first case it is turned back according to fixed rules, while in the second there is no uniformity about the reflection. The page on which this is printed appears to be white because, owing to the roughness of the paper, of the irregular reflection of the light which falls upon it. Or, if we powder a sheet of glass, the powder seems to be white for a similar reason; there are many surfaces formed from which irregular reflection takes place.

Regular Reflection of Light.—Light is regularly reflected from a plane mirror, that is, a flat reflecting surface. Such a mirror can be made from a variety of substances, but the most common is bright metal or silvered glass. To find the relation between the angle at which a beam strikes a mirror and the angle at which it is reflected, the following experiment should be performed.

EXPT. 162.—Fix two slabs of wood at right angles as in Fig. 73, AB, CD. Against the upright slab place a piece of glass EF with blackened back so that reflection only takes place from the front. Upon the horizontal slab place a sheet of white paper. Stick a pin *b* in the wood against the glass, and place another pin near the position *a*. Now procure another pin and stick it into the wood at *c* in such a position that *c*, *b*, and the image of *a* are in a straight line. Draw with a finely pointed pencil a line along the edge of the glass *xy*; then take glass and pins away.

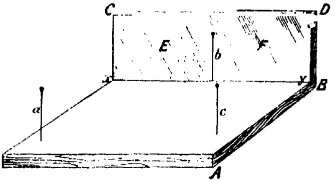


FIG. 73.—Arrangement to illustrate the Law of Reflection.

The paper will be marked by the pin-holes and the line *xy*. Draw lines through the pin-holes, and at *b* a normal to *xy*, that is, a line perpendicular to *xy*. Measure the angles *abd*, *cbd*, and compare them (Fig. 74). Repeat the experiment two or three times, with the pins in different positions, and so determine that the angle of incidence and the angle of reflection are equal.

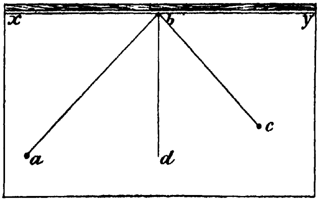


FIG. 74.—Angles of Incidence and Reflection.

The same facts can be demonstrated and shown to the whole class in another way :—

> EXPT. 163.—Fasten a little whitened, wooden stick with wax perpendicularly at the centre of a plane looking-glass. Cast upon the mirror, at the foot of the rod, a beam of parallel rays from the lantern, or a sunbeam coming through a hole in a screen. Notice (*a*) that the reflected beam always makes the same angles with the mirror and the stick as the incident beam does, and (*b*) that the incident beam, the stick, and the reflected beam all lie in one plane (Fig. 75).

From these observations we learn that the light strikes the mirror at a certain angle and leaves it at the same angle. The angle at which the light or any sort of wave strikes the reflecting surface is called the **angle of incidence**, and the wave an **incident wave.** The angle at which the wave leaves this surface is known as the **angle of reflection**, and the wave as it leaves the **reflected wave.**

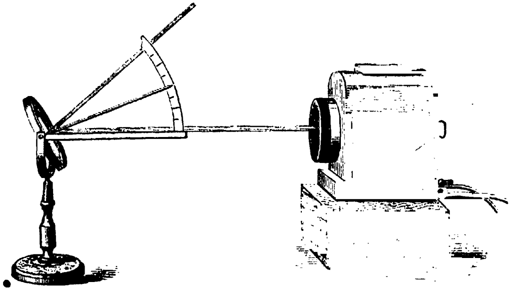


FIG. 75.—The Laws of the Reflection of Light.

Laws of Regular Reflection.—From the experiments and considerations which have just been described we arrive at the two laws which must be very carefully remembered :—

1. **The line representing the reflected wave is in the same plane with the normal and the line representing the incident wave, and is on the opposite side of the normal from the incident line.**

2. **The angle of incidence is equal to the angle of reflection.**

The student must bear in mind that these laws apply equally to all forms of radiation, whether in the ether or in air; *i.e.*, to all ether waves and to all sound waves.

It was also learnt from Expt. 162 that when a wave strikes a reflecting surface *normally*, *i.e.*, having travelled along the normal, it is reflected back along the same line.

Formation of an Image by a Plane Mirror.—These two rules enable the formation of an image by a plane mirror to be easily understood.

Let MM be the plane mirror, and A a bright spot of light like the head of the pin in Expt. 162. First see what happens to the light ray which leaves A and strikes the mirror normally. It is reflected back along the same line and the reflected ray appears to come from a point A′ along A′A. In the case of any other rays, such as AB, the light ray is reflected in such a way that the angle of reflection CBD is equal to the angle of incidence ABC and appears to an eye placed as in Fig. 76 to come along BD from a point A′, where BD produced meets AA′. If we make the same construction for any other ray AB′ it will be reflected and appear to be coming along B′D′, which produced backwards will pass through the same point A′. A′ is therefore the image of A, and it can be easily proved by geometry that A′ is as far behind the mirror as A is in front of it.

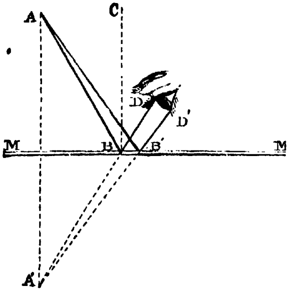


FIG. 76.—The Formation of an Image by a Plane Mirror.

Reflection from Spherical Mirrors.—A spherical mirror is a part of a spherical surface which has the power of reflecting. It may be either *concave* or *convex*, the former if the reflection takes place from the hollow side, the latter if from the bulging side. The centre of the sphere of which it forms part will evidently be the centre from which the part of it constituting the mirror was struck, and this point is called the *centre of* *curvature.* The distance from this point to the reflecting surface is the *radius of curvature.* Thus in Fig. 77 *c* is the centre of curvature and *c*M, *cd*, *c*M′, are all radii of curvature. MM′ is called the diameter or aperture of the mirror and *d* is called by many names, perhaps *pole* of the mirror is the best. A line going through the pole and centre of curvature is the *principal axis* of the mirror, any other radius produced being a *secondary axis.* We know from geometry that every radius is at right angles to the tangent at the point where it cuts the circle, and since we can consider the tangent and circle as coincident at this point; from what has been already said it will be clear that the radii are normals to the mirror. Evidently, then, if we place a luminous object at the centre of curvature we shall have all the rays of light from it reflected back along the lines of incidence, or the image will be formed at the same place as the object.

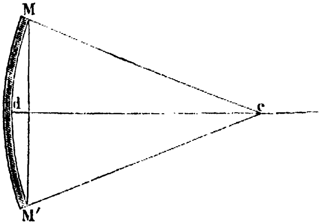


FIG. 77.—Reflection from a Spherical Mirror.

EXPT. 164.—Procure a concave mirror and cover it with black paper, except a small part at the centre or round the pole. That is, let the aperture of the mirror be small. Allow rays of sunlight to fall upon it (these come from so great a distance that they can be considered *parallel*). Move a very small paper screen up and down in front of the reflecting surface so as not to cut off the incident rays. Notice that at a certain point a clear image of the sun is formed, and probably the screen will be burnt.

The point so obtained is called the **principal focus** of the mirror. In Fig. 78, F represents this point and C the centre of curvature. The parallel lines show the direction of the sun's rays. This point F is midway between the pole and C, or the *focal length* is half the radius of curvature.

Refraction of Light.—Up to the present the light rays have been supposed to be moving through a medium of a uniform density throughout, such as is the case when it travels through air. When this is so, as we have seen, light travels in straight lines, and, coming into contact with a reflecting medium, is turned back, according to the laws which have been studied.

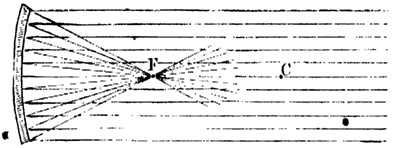


FIG. 78.—The Principal Focus of a Spherical Mirror.

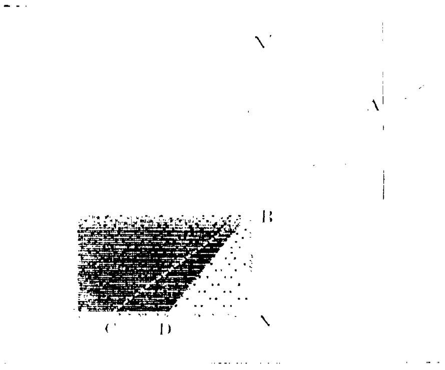


FIG. 79.—Refraction of Light.

If, however, light passes from one medium into another of a different density the propagation of the wave is no longer rectilinear, the passage from one medium into the other is accompanied by a bending of its path. This bending is known as

refraction, and the ray is said to be **refracted.** We will describe how this can be shown experimentally in the case of light-waves, but what holds good of waves of this particular length holds true of all ether-waves, as well as of any other kind of radiation.

EXPT. 165.—Procure a rectangular metal box, such as a cigarette box, and put a wooden or metal scale on the bottom. In a darkened room let sunlight fall slantwise against the edge. The side of the box throws a shadow which reaches, say, to C (Fig. 79), which, since light travels in straight lines in the same medium, is a continuation of the rays of sunlight A B. Without disturbing anything fill the box with water. The shadow no longer reaches to C, but only as far as D. Clearly the light-waves have been bent or refracted out of their original course. Notice that NN′ is the normal, and that the light travelling from the comparatively rare air into the comparatively dense water is refracted towards the normal.

This experiment illustrates the general fact of the deviation experienced by a ray of light in passing from one medium to another of different density. The **laws of refraction** can be determined by another simple experiment.

EXPT. 166.—Upon a piece of board, ABCD (Fig. 80), place sheet of paper, and upon the paper put a piece of fairly thick glass EF with parallel sides (a thick piece of glass from a box of weights, a paper-weight, or a number of slips of microscope glass will do very well). Rule along the edges of the glass with a finely-pointed pencil. Place two pins *a*, *b* as shown in the illustration, and then, *looking through the glass* from the other side, stick in the pins *c*, *d* so that all four appear in a straight line.

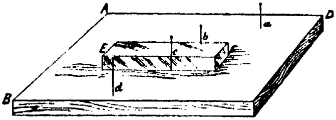


FIG. 80.—Method of Determining Refraction of Light.

Now take away the glass and pins and join the pin-holes on the paper as shown in Fig. 81. Draw the normal *ebf*, and the circle *elfg*. Draw *lm* and *gk* perpendicular to *ebf*, and compare the lengths *lm*, *gk*. Obtain the ratio $\frac{lm}{gk}$ for different positions of the pins; it will be found practically the same in all cases in which the same material is used.

Notice that the direction of the ray *cd*, emerging from the glass, is parallel to that of *ab*.

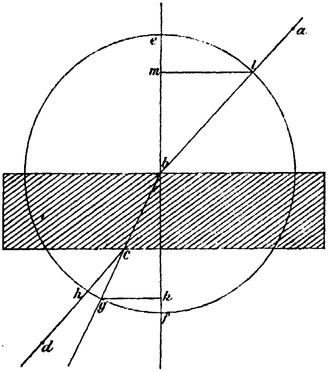


FIG. 81.—Refraction of Light passing through Glass.

These experiments will enable the general rules of refraction to be well understood.

Rules of Refraction.—In Fig. 82 the shaded lower part of the diagram represents a denser medium than the unshaded upper portion. The word denser is used here, and in similar connections, to mean optically denser, and must not be confused with what has been said of the density of bodies in Chapter VII. Let RI represent a ray passing from the rarer to the denser medium, or the ray incident on the surface of the denser medium at I. The angle RI makes with the normal at I is the angle of incidence. The ray is bent, as we have learnt; instead of continuing its course in a straight line along IR′ it is refracted and travels in the direction of IS which represents the refracted ray, the angle SIP being the angle of refraction. The angle R′IS which represents the amount the ray has been turned out of its original path is called the *angle of deviation.* With the centre I and any convenient radius describe a circle, and from the points where it cuts the incident and refracted rays draw perpendiculars on to the normal as in the figure. Also drop a perpendicular from the point R′. It is clear from geometry that R′P′ is equal to the perpendicular let fall from the point where the incident ray cuts the circle on to the normal. The ratio between the lengths of R′P′ and SP is constant for the same two media, *e.g.*, air and water, whatever the angle of incidence. This ratio is called the *index of refraction.* Its value for air and water is about $\frac{4}{3}$; for air and glass roughly $\frac{3}{2}$ depending upon the kind of glass.

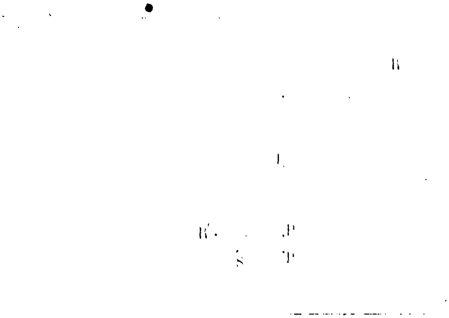


FIG. 82.—To explain the Rules of Refraction.

The laws of refraction are then :—

1. **The incident ray, the normal, and the refracted ray are all in the same plane. The incident and refracted rays are on opposite sides of the normal.**

2. **If a circle be described about the point of incidence, and perpendiculars be dropped upon the normal from the intersections of this circle with the incident and refracted rays, the ratio of the lengths of these perpendiculars is constant for any two given media.**

Effects of Refraction.—The following experiments illustrate further the phenomena and effects of refraction :—

EXPT. 107.—Place a bright object, say a coin, on the bottom of an empty basin, and arrange your eye so that the object is *just* hidden by the edge of the basin. Get somebody to pour water into the basin. You will now be able to see the coin without any movement of your eye having taken place. Evidently there has been some bending of the direction of the light rays somewhere.

In Fig. 83 let C be the position of the coin in which it is just hidden, as far as the eye is concerned, by the edge of the empty basin. If the rays from the coin C be continued in straight lines these lines will evidently pass above the eye. Now when the water is put in, these rays, which before miss the eye, are refracted in passing out of the water and just enter the eye, making the coin appear to be in the position C′. The right side of the vessel, if continued upwards, represents the normal, and evidently in passing out of water into air the light-waves are bent away from the normal.

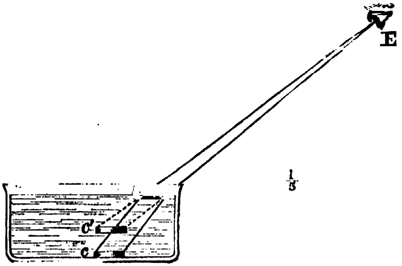


FIG. 83.—Coin Experiment to show Refraction.

EXPT. 168.—Stretch a wire across the opening in the cap of a lantern. Focus the shadow of the wire on a screen. Interpose a strip of glass in front of the wire and tilt the strip. Notice the displacement of part of the shadow.

EXPT. 169.—Place a glass cell before the lantern and focus it on the screen. Let the surface of the water be visible. Put a lump of ice on the water, and observe the streaky appearance of the illuminated part of the screen. Add syrup, alcohol, and hot water by a pipette. Notice similar effects.

Pour the vapour of ether out of a bottle, and observe the appearance on the screen. Examine the shadows of burning gas, a red-hot poker, or of platinum wire through which a current of electricity is passing.

It is the refraction of light in its passage from one medium into another of different density which explains several very familiar observations. A stick held in anything other than a perpendicular position in water appears to be bent upwards (Fig. 84). Standing bodies of water always appear shallower than they really are on account of refraction.

Expt. 170.—Fill a glass cylinder (Fig. 85) with water, and place a coin at the bottom. On looking straight down through the water the coin appears nearer the surface than it really is.

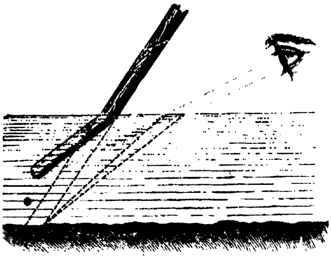


Fig. 84.—To illustrate why a Stick placed slantingly in Water appears to be bent upwards.

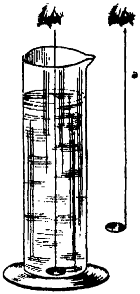


Fig. 85.—To illustrate an effect of Refraction.

Hold another coin near the outside of the cylinder and place it at such a height that the two coins appear at the same level. The amount by which the coin in the water is apparently elevated by refraction can thus be found. The length of the column of water through which the coin is observed, divided by the distance from the top of the water to the outside coin, gives the index of refraction of water.

Repeat the experiment with methylated spirit.

Refraction through a Prism.—In optics a prism is defined as "a wedge-shaped piece of glass contained between two planes called its faces, which intersect in a line termed the edge of the prism." It may also be defined as a solid having three rectangular sides (Fig. 86).

> **Expt. 171.**—Focus on the screen a small circular hole in the cap of the lantern. Introduce a wedge of glass, and show that the spot is moved towards the base of the wedge. Let the first face of the wedge be oblique to the beam. Put a second wedge with the first, face to face and base to base, so as to form a single wedge of double the angle. Notice the increased displacement of the light spot. Put them edge to base, notice the emergent beam is parallel to the incident.

The amount of deviation of the beam thus depends upon the angle of the wedge or prism; it also depends upon the material of which the prism is made and the nature of the incident light.

Path of a Ray of Light through a Prism.—In Fig. 87 let the triangle *abc* represent a section of the prism at right angles to its faces, such as we should see by looking at the end of it. Suppose DE is a ray of light striking the face *ab* of the prism. The light on entering the prism passes from the air into the glass, or *from a rarer into a denser medium*, and is, as was seen in Expt. 166, bent *towards* a line drawn perpendicular to the face of the prism at the point where the ray of light strikes it. It consequently travels along the line EE′ until it reaches the face *ac* of the prism. Here it passes from the glass into the air, *i.e.*, *from a denser into a rarer medium*, and is, in such circumstances, bent *from* the perpendicular, and travels along the line *E′D′*. In every such passage through a prism it is

FIG. 86.—A Prism.

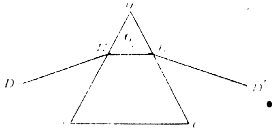


FIG. 87.—Path of a Ray of Light through a Prism.

noticed that the light is always bent or refracted towards the thick part of the prism.

The course of a ray of light through a prism can be found by a simple experiment.

EXPT. 172.—Stand a prism upright, that is, upon one of its ends, upon a piece of white paper. Stick two pins into the paper in the positions D E (Fig. 87), place two more E′ D′ on the opposite side of the prism so that the four appear in a straight line when looking through the prism. Draw the outline of the prism *abc*, and then take away the prism and the pins and connect the pinholes as shown in the diagram. It will be found that the ray is bent towards the base of the prism both when it enters and emerges.

Refraction through a Lens.—Most lenses are of glass with curved surfaces which are portions of spheres. In some lenses one surface is quite plane. All lenses can be divided into two classes—*convex* or *converging*, and *concave* or *diverging*. Converging lenses can be told at once by their power of forming an image of a distant object like the sun, or by that of magnifying. Concave lenses form no image in this way, and, moreover, instead of magnifying they make objects appear smaller when viewed through them.

To understand their action upon the course of rays of light through them it is simplest to regard them as being built up of parts of prisms in contact, as shown in Fig 88, where a convex lens is built up in this way. A ray of light falling upon any one of these prisms is refracted towards its thicker part, and consequently they all converge towards a point, which, if the incident rays are parallel, is known as the *principal focus* of the lens, as F in Fig. 89. To actually find the distance of the principal focus away from the centre of the lens, that is, its *focal length*, it is only necessary to form an image of the sun by it on a screen and to measure the distance between the lens and the screen.

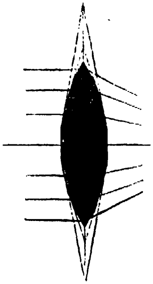


FIG. 88.—Lens built up of parts of Prisms.

Refraction is accompanied by Dispersion.—In all the cases of refraction which have been considered the phenomena have been described as if all the ether-waves, which are contained in white sunlight, are bent equally, but this is really not so. If the experiments are performed as described, in every case it will be noticed that the image formed by the refracted ray is coloured round its edges. To prevent this complication, and to make the descriptions quite correct, we must suppose that waves of a given length are used, or, as it is called, *monochromatic light*, such, for example, as could be obtained by burning methylated spirit in which common salt is dissolved.

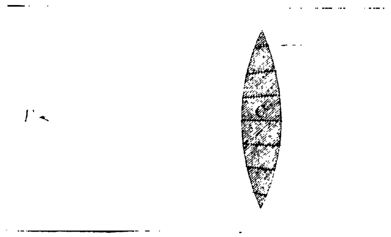


FIG. 89.—To show Principal Focus of a Double Convex Lens.

The shortest, most rapid waves are bent most; the slowest, longest waves are bent least. This can be very prettily seen by performing an experiment which was first done by Newton; improving on his method, however, by using a slit instead of a round hole, through which to let the light pass.

The Analysis of Light by a Prism.—EXPT. 173.—Allow sunlight to fall through a slit or hole, Fig. 90, into a darkened room. Just inside the room arrange a prism as shown. The component wave-lengths of the white light will be differently refracted by the prism, the rapid vibrations of the ether which give rise to the sensation of violet will be most bent, the slowest vibrations or the red waves will be bent least, and a band of colour, continuous between these limits and comprised of the colours of the rainbow, will be formed.

In this experiment the light is said to be **dispersed** owing to the different *refrangibility* of the various kinds of light. An examination of the band of colour or **spectrum** will show that one colour shades imperceptibly into the next. There is, then, an infinite number of waves of different lengths comprised in the white light from the sun, and each ray is bent to an extent depending on its wave-length.

If the decomposed sunlight, instead of being collected on a screen, be passed through a second prism similarly arranged, it will be seen that the band is longer or the dispersion is greater. The amount of dispersion also depends upon the material of which the prism is made. Glass produces a much greater amount of dispersion than water ; flint-glass possesses twice the dispersive power of crown glass ; carbon bisulphide, again, has more dispersive power even than flint-glass.

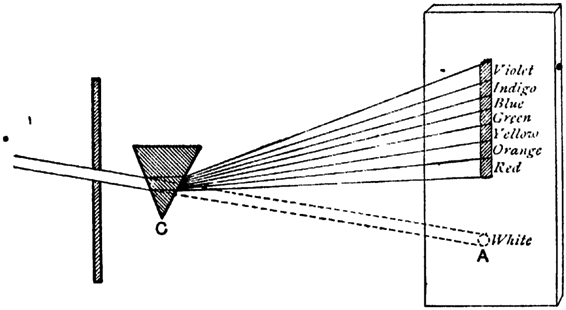


FIG. 90.—The Analysis of Light by a Prism.

Although a continuous band of colour is observed when sunlight, or limelight, or a gas or candle flame is seen through a prism, this continuous spectrum is not always produced. If when substances such as sodium, strontium, and lithium, or their compounds, are burnt in a non-luminous flame, and observed through a prism a discontinuous spectrum is seen. The following experiment, described by Sir Norman Lockyer, illustrates this.

EXPT. 174.—Procure a strip of wood about 12 × 1 × ½ inch, and glue a glass prism upright at one end of it. Stick a pin or a needle upright in the other end. Allow the needle to be illuminated by the flame of a spirit lamp into which salt is gradually allowed to fall. An orange-coloured image of the needle is seen. Now illuminate the needle by a candle or gas flame, taking care that the direct light from the candle does not fall upon the face of the prism (Fig. 91). No longer is a single image of the needle seen, but a complete band of colour from red to blue. This is produced by an innumerable multitude of images close together.

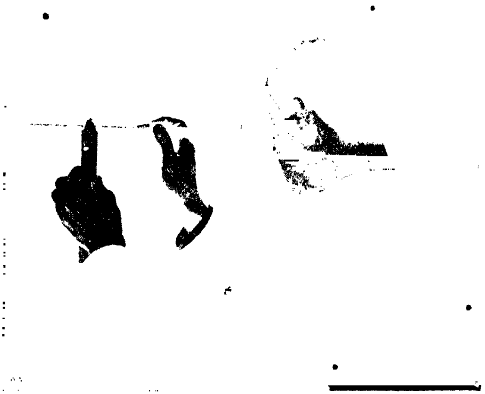


FIG. 91.—To illustrate Experiment 174.

A prism may thus be used, and is used, to analyse light. The light of incandescent sodium vapour when observed through a prism is characterised by the orange-coloured ray seen in the above experiment with common salt, which contains sodium, and the light emitted by other substances when burning are each distinguished by rays of a particular colour and position in the spectrum.

Recomposition of White Light.—If in Expt. 173, after having decomposed sunlight to form a spectrum, we place a second similar prism after the first *with its refracting angle reversed*, it will be seen that the spectrum disappears, and in its place we get a white image of the slit, not opposite the slit, but displaced owing to the refraction produced by the prisms. But

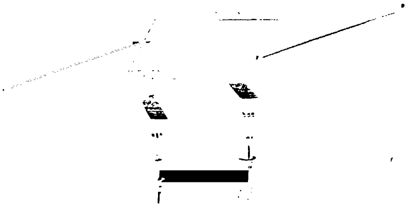


FIG. 92.—Recomposition of Light by a Second Prism.

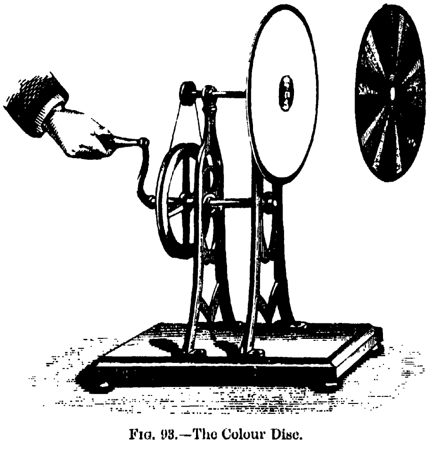


FIG. 93.—The Colour Disc.

the refraction is unaccompanied by dispersion. The second prism undoes the dispersive work of the first (Fig. 92).

Or if, after the first prism, a double convex lens were arranged so that the decomposed light falls upon it, it will be found that the dispersed light is recombined and a white image is formed at the principal focus of the lens.

The Colour Disc.—This is an arrangement of Newton's for demonstrating that white light can be made by combining the various colours of the spectrum.

> EXPT. 175.—Upon a round piece of card paint sectors of the different colours contained in the spectrum, arranging the areas of the coloured sectors as nearly as possible in the proportion in which they occur in the spectrum.
>
> Place the card upon a whirling table (Fig. 93) or upon a top, and rotate it rapidly, when it will be found that light from the card gives rise to the sensation of white or gray.

CHIEF POINTS OF CHAPTER XI.

Light, like every kind of radiation, is a form of energy. Hence it is a process of transference of energy by ether-waves. These ether-waves which affect the retina are known as light.

Rectilinear Propagation of Light.—Light travels in straight lines when propagated through any one medium, but often has its direction changed when passing from one medium to another (*see* Refraction).

Reflection.—Light in being reflected from suitable surfaces, obeys the following laws:—

(1) The reflected ray lies in the same plane as the incident ray and the normal.

(2) The angle of reflection is equal to the angle of incidence.

Reflection from Spherical Mirrors.—The centre of the sphere of which the mirror is a part is called the *centre of curvature.* The distance from this point to the reflecting surface is called the *radius of curvature.* The point to which all parallel rays converge after reflection is called the *principal focus.* The distance of the principal focus from the *pole* of the mirror is called the *focal length of the mirror* and is equal to one half of its radius of curvature.

Refraction of Light.—A ray of light passing from a less dense to a more dense medium is bent towards the perpendicular to the separating surface and conversely. The laws of refraction can be stated thus:—

(1) The incident ray, the normal, and the refracted ray are all in the same plane. The incident and refracted rays are on opposite sides of the normal.

(2) If a circle be described about the point of incidence and perpendiculars be dropped upon the normal, from the intersections of this circle with the incident and refracted rays, the ratio of the lengths of these perpendiculars is constant for any two given media.

Refraction by a Prism.—The deviation of a beam of light caused by the action of a prism depends upon (*a*) the angle of the prism, (*b*) its material, (*c*) the thickness of it traversed, (*d*) the length of the ether-waves.

Refraction through a Lens.—Most lenses are of glass with curved surfaces which are portions of spheres. They can be divided into convex or converging and concave or diverging lenses. The point to which parallel rays after refraction converge, or appear to diverge, is called the *principal focus*.

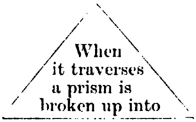
Analysis of Light by a Prism.—The points to be remembered in connection with this are :—

(1) That if a beam of monochromatic light, that is, light of one wave length, traverses a prism, it is bent out of its original direction, but the amount of bending produced by any particular prism depends upon the colour of the beam used, being greatest for violet light and least for red light.

(2) That if light from any source pass through a prism, it is broken up or analysed into its different components, each ray of elementary colour that enters into the composition of the light being bent by a different amount.

(3) That if a beam of white light from a slit pass through a prism, it emerges as a coloured strip, termed a *spectrum*.

A beam of white light,



When
it traverses
a prism is
broken up into

Red
Orange
Yellow
Green
Blue
Indigo
Violet

White light can, therefore, like a chemical compound, be broken up into simpler elements.

Recomposition of Light—The Colour Disc.—The analysis of white light is noted above; the synthesis, or recomposition, can be effected—

(1) By making the coloured band or spectrum, produced when light has passed through one prism, traverse a second prism having its refracting angle reversed.

(2) By rotating a disc of cardboard painted in segments of violet, blue, green, yellow, orange, and red,

QUESTIONS ON CHAPTER XI.

1. How could you show that light travels in straight lines?

2. Describe experiments to illustrate the characteristics of transverse and longitudinal waves.

3. State the laws of the reflection of light.

4. In what direction is a beam of light deviated when it passes from a less dense to a more dense medium?

5. Name some common phenomena produced by the refraction of light.

6. Draw a triangle to represent a section of a prism and show how a beam of light is bent by passing through the prism.

7. State how light can be experimentally analysed by a prism.

8. How would you prove that white light is composed of several colours, and how can these colours be combined by experiment to form white light?

9. Explain the statement "Light is a form of energy."

10. Describe a simple experiment to prove the laws of reflection.

11. What experiment could you perform to explain such every-day phenomena as the shallow appearance of a brook, and the apparent bending of a stick which is slantingly plunged into water?

12. Explain the terms *principal focus* and *radius of curvature* as applied to a spherical mirror.

13. Describe some commonly observed effects of the refraction of light.

14. What is meant by the *Dispersion of Light?* On what fact does it depend?

15. Explain the term *refrangibility* as applied to a ray of light. Are rays of all colours equally refrangible?

16. A bright bead is placed at the bottom of a basin of water, and a person stands in such a position that he can see it over the edge of the basin. While he is looking, the water is drawn off. How will this affect his view? Draw a diagram showing the direction of a ray of light passing from the bead through the water and the air in each case.

17. How can it be proved that:—

(*a*) White light is a mixture of many colours. (*b*) Different colours have different degrees of refrangibility.

18. A room is lighted by one small window which can be completely covered by a screen of dark red glass, or by one of blue glass, or by both put together. In the room is a stick of sealing-wax, a piece of blue ribbon and a lily. What colour does each of these things appear when (*a*) the red screen covers the window, (*b*) the blue screen covers the window, (*c*) when both screens cover the window?

19. Explain why if a sheet of paper be placed behind a pin-hole in a thick sheet of cardboard, an image of a brilliantly illuminated object on the other side of the cardboard will be formed on the paper. Why is the image fainter, but more clearly defined if the pin-hole is very small?

20. A thick layer of a transparent liquid floats on the surface of water. Trace the course of a ray of light from an object immersed in the water through the floating liquid to the air.

21. Describe and explain the effects observed when cards coloured bright red, green, and blue respectively are passed from the red to the blue end of the spectrum.

22. Describe an arrangement by means of which a spectrum may be formed upon a screen.

If the light is made to fall upon a piece of red glass before reaching the screen, how and why will the spectrum be affected? What would the effect have been if blue glass had been used?

CHAPTER XII

THE CONSTRUCTION OF SIMPLE VOLTAIC CELLS

THE facts to be described in this chapter and the next refer to the production of electric currents by means of simple voltaic cells, and to chemical changes which such electric currents are able to produce. In order to obtain a clear idea of the use of the various materials employed in the construction of voltaic cells, and the chemical effects to which electric currents are able to give rise, it is necessary to know something about the principles of chemistry. This chapter and the following one will, therefore, be more profitably read if they are taken after the section on chemistry than if they are studied at this stage.

Fundamental Experiments.—EXPT. 176.—Prepare some dilute sulphuric acid containing one part of strong acid to eight parts of water. First measure out the water into a large beaker, and then gently pour the measured quantity of strong acid into the water, keeping the latter briskly stirred with a glass stirring rod. Having noticed the large amount of heat generated set the mixture on one side to cool.

EXPT. 177.—Plunge a strip of ordinary commercial zinc into a beaker of cold dilute sulphuric acid prepared as in the last experiment. Notice the brisk evolution of gas which takes place. What gas is given off? (*See* p. 189.)

EXPT. 178.—Repeat Expt. 177, substituting first a rod of pure zinc and then a strip of copper. Observe that there is no chemical action (p. 189) in either case.

EXPT. 179.—Place the rod of pure zinc and the strip of copper into the dilute acid, taking care that the two metals do not touch one another. No gas is given off from either metal.

Expt. 180.—Now tilt the pieces of metal towards one another until they touch outside the liquid, or connect them with a wire as in Fig. 94. Observe that bubbles of gas appear on the copper plate.

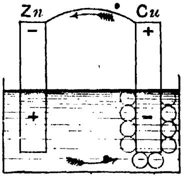


Fig. 94.—To illustrate Experiments 179 and 180.

Expt. 181.—Take the pieces of copper and zinc in the last experiment out of the acid, wash them under the tap, and carefully dry them on clean blotting paper. Weigh both pieces of metal and record their weights thus:—

Weight of rod of pure zinc — grams.
Weight of strip of copper — grams.

Now replace them in the dilute acid and make them touch outside the liquid, or connect them by a piece of copper wire, and allow the action which ensues to continue for ten or fifteen minutes. Then remove them from the acid and wash, dry and weigh them as before. Record your result beneath the last entry thus:—

After 15 minutes' action.
Weight of zinc rod ... ... — grams.
Weight of copper strip ... ... — grams.

Notice that while the zinc has diminished in weight the weight of the copper has remained unaltered.

Expt. 182.—Prepare a plate of amalgamated zinc by dipping a plate of ordinary commercial zinc into dilute sulphuric acid, and after it has been acted upon for a minute or two rub some mercury completely over its surface with a piece of cloth. Repeat Expt. 178, and observe that amalgamated zinc behaves just like pure zinc.

Properties of the Wire connecting the Plates of Copper and Zinc.—Expt. 183.—Into some dilute sulphuric acid contained in a beaker plunge a plate of amalgamated zinc and one of copper, to each of which a copper wire is attached by a suitable binding screw. There is no chemical action until the free ends of the wires are joined, when, as before, bubbles of gas are seen to collect on the copper plate.

Expt. 184.—Procure an ordinary compass needle, which consists of a light magnet, suitably supported at its middle point, so that it can move freely in a horizontal plane. Bring this up and try and arrange matters so that the wire of Expt. 183 in connection with the copper and zinc plates is parallel to the magnet when the latter is pointing north and south (Fig. 95). Notice that this is impossible. The wire exerts some sort of power on the needle.

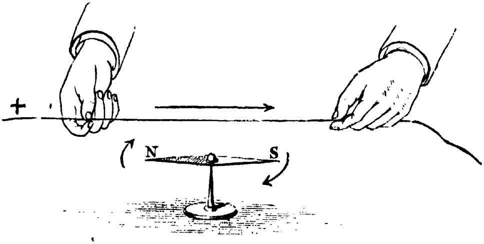


FIG. 95.—Experiment 184.

EXPT. 185.—Wind the wire which is connected with the copper and zinc plates round a piece of soft iron, as in Fig. 96. Notice that the piece of soft iron will attract iron filings.

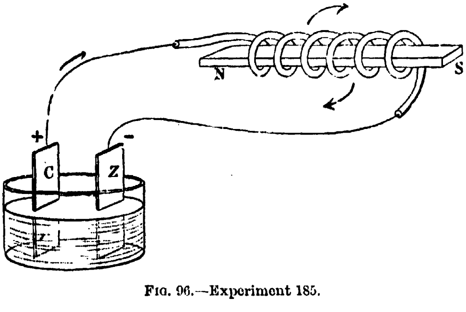


FIG. 96.—Experiment 185.

EXPT. 186.—Repeat Expt. 184, and notice that after a time the power which the wire exerts over the magnetic needle becomes feebler, and by and by ceases. Rub the copper plate with a piece of wood until all the bubbles of gas which had collected have disappeared, and notice that the power of the wire to deflect the magnetic needle is regained.

What the Experiments Teach.—The student should have no difficulty in understanding now that a wire which connects a strip of zinc and a strip of copper, which are dipping into dilute sulphuric acid, has new properties. It is in a different condition from an ordinary piece of such wire. Its new condition is described by saying that *an electric current is passing along it.* The arrangement is, in fact, a means of developing an **electric current**, and is called a **voltaic cell.**

Moreover, the use of such words as "flow" and "current" will probably suggest previous facts which the reader has learnt. He has seen that a flow of heat takes place from a body of high temperature to one of low temperature when they are placed in contact, and that such flow continues until both bodies are of the same temperature.

Similarly, there is a flow of water from one vessel to another, which are in connection, if the level of the water in one vessel is higher than in the other. Hence we are face to face with the question, What difference is there between the copper and the zinc which causes the condition of things called the electric current? The name given to the difference of condition in the copper and zinc which corresponds to temperature and water-level is **difference of potential.** The electric current continues to flow along the copper wire until the potential of the two plates becomes the same, when it ceases. This is the state of affairs after the bubbles of the gas, hydrogen, have collected on the copper plate.

The electrical condition of the plates of copper and zinc becomes altered by placing them in the sulphuric acid, but the final state is different in the two cases. The copper plate is at a higher potential than the zinc plate, and, consequently, when the plates are joined outside the liquid by a copper wire there is a flow of the electric current *from the copper to the zinc.* The difference of potential which causes the flow is maintained by the solution of the zinc in the acid. Or, speaking in terms of energy (Chap. IX.), the work of maintaining the current is performed by the solution of the zinc. This is similar to the maintenance of the work done by a steam-engine by the burning of the coal in the furnace. It was for this reason that we found the zinc weighed less in Expt. 181, after the electric current had flowed for some time along the joined wires.

It is customary to give the name **electro-motive force** to the difference of potential which we have seen exists between the plates of copper and zinc. As the student will learn when he continues his study of the electric current, this difference of potential, or electro-motive force, can be measured by means of an instrument devised by Lord Kelvin, called the **Quadrant Electrometer.**[1] But though the term electro-motive force is used to speak of this difference of potential, it is not a force in the sense in which the word was used in Chapter V. A force is that which causes motion in matter, but an electro-motive force results only in a motion of electricity, which is not matter at all.

Voltaic Cells.—The simplest arrangement for developing an electric current by means of chemical action has been now described, but there are many others which are much better. They are all called "voltaic" cells after the name of the physicist **Volta,** who was the first to develop an electric current by chemical means.

To be of any practical value the voltaic cell must not cease in a short time to give a current as the simple device adopted in Expt. 186 did, as soon as bubbles of gas had collected over the copper plates. Some means have to be taken to prevent this stopping of the current by the bubbles of hydrogen, or, as it is called, the **polarisation** of the cell.

Polarisation is prevented in two ways: (1) by mechanical, (2) by chemical means.

Prevention of Polarisation by Mechanical Means.—Smee's Cell.—Smee substituted a plate of silver on which platinum had been deposited, that is, a plate of *platinised silver*, for the copper plate in Volta's simple cell. The platinum on the silver plate causes it to be very rough, with the result that there is a continual tendency for the bubbles of gas to rise from the points of roughness, up through the liquid, and to escape at the surface.

In other patterns of cell a mechanical device for keeping the liquid stirred near the copper plate is adopted, the agitation of the liquid being quite enough to prevent the collection of gas bubbles.

Chemical Changes in a Simple Cell.—To understand clearly the principle of the construction of cells in which the difficulty of polarisation is got over by chemical means, it will be necessary to explain the chemical changes occurring in the simple cell already described. The facts we have learnt are—first, that the zinc gradually dissolves in the acid, and second, bubbles of hydrogen continually collect on the copper plate. Since we have proved by experiment that dilute sulphuric acid has no action on copper, it is evident that though the bubbles of hydrogen collect on the copper plate they were not turned out of the sulphuric acid by the copper. As we shall learn in the chemical section of this book, the chemical change which takes place when ordinary commercial zinc is placed in sulphuric acid can be represented by an equation:—

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[1] [illegible] Prof. S. P. Thompson's *Electricity and Magnetism*, p. 272.

When SULPHURIC ACID acts upon ZINC we have ZINC SULPHATE formed
and HYDROGEN given off;

or more shortly by chemical symbols,

$$H_2SO_4 + Zn = ZnSO_4 + H_2.$$

The hydrogen is, then, really first liberated near the piece of zinc, and it is supposed that at the moment of its liberation it breaks up the adjoining molecule of sulphuric acid, taking the place of the hydrogen therein, and turning out that which was already combined with the other elements in the molecule of sulphuric acid.

This can be shown a little more clearly, perhaps, by using the symbol H′ for the hydrogen first turned out by the zinc, and H″ for that in the neighbouring molecule of sulphuric acid of which we have spoken. Then the interchange of hydrogen in this molecule of sulphuric acid can be expressed by an equation:—

$$H''_2SO_4 + H'_2 = H'_2SO_4 + H''_2.$$

This interchange of the hydrogen goes on right through the liquid until the copper plate is reached. We have the same reaction going on an almost infinite number of times:—

$$H'''_2SO_4 + H''_2 = H''_2SO_4 + H'''_2$$

until finally there is no longer a molecule of sulphuric acid to be decomposed, and the free hydrogen is deposited on the copper plate. It is customary to represent the state of affairs in the cell graphically, and since it is very important that a proper conception of the continuous decomposition of sulphuric acid molecules should be formed, we also give this method, as it may help some readers better than the previous plan. In the illustration (Fig. 97) the second row of oval figures, stretching between the zinc and copper plates, A and B respectively, is to represent the condition of the molecules of the acid before the plates are joined by a wire. The shaded part of each figure represents the hydrogen in the molecule, and the unshaded portion the group of elements, SO_4, which complete the constitution of the sulphuric acid molecule. The third row shows the result of the series of decompositions which take place when the plates are joined by a wire. At the zinc plate A zinc sulphate is shown formed, and at the copper plate B free hydrogen is represented.[1]

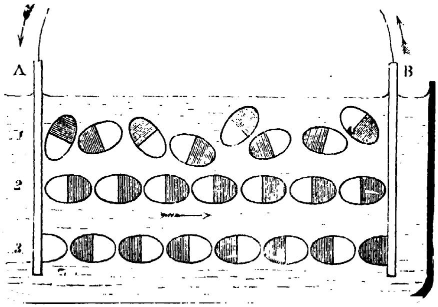

FIG. 97.—The passage of an Electric Current in a Simple Cell.

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[1] The student must, however, in view of more recent work on the subject, avoid picturing the process as taking place with the regularity indicated in the above explanation and in Fig. 97. It is now believed, that most probably, the solution of acid contains not only H_2SO_4 molecules, but also uncombined hydrogen atoms (or as we shall learn to call them *ions*) and uncombined SO_4 groups. These ions are thought to be electrically charged, those of hydrogen positively and the SO_4 groups negatively. Consequently, whenever these separate ions combine, the resulting H_2SO_4 molecule is electrically neutral. Each molecule of H_2SO_4 is being constantly decomposed into its ions, and these separated ions are constantly recombining in other cases to form new H_2SO_4. The net result of these decompositions and recompositions is that the ratio of decomposed to undecomposed molecules remains constant. The result of the union of the zinc with the negatively electrified SO_4 ions is to leave it negatively electrified. The preponderance of positively electrified hydrogen ions in the liquid causes it to be positively electrified. Or, using the word "potential," which the student has learnt, the zinc is left at a lower potential than the liquid. But the zinc is in metallic connection with the copper, which must also be at a lower potential than the liquid. The positively electrified hydrogen ions hence travel towards the copper plate, where *after giving up* their positive charges they escape as *gaseous hydrogen*. By this means the current is maintained.

We should, perhaps, be more accurate if we regarded the negative electrification of the zinc as due, not to the union with it of the negative SO_4 ions, but to the departure from it, into the solution, of positive zinc ions.

Prevention of Polarisation by Chemical Means—1. **Daniell's Cell.**—*The deposition of copper instead of hydrogen.* In most of the cells in which polarisation is prevented by chemical means there are really two cells, the one placed inside the other. The inner one is made of some porous kind of earthenware which permits a slow passage through it of the liquids on either side of it. In Daniell's cell the outer vessel is of copper and serves as the copper plate. The outer vessel contains a solution of copper sulphate (blue vitriol), the strength of which is maintained by placing some crystals of the same substance on a tray, which extends round the top of the inside of the copper vessel (Fig. 98). The inner porous pot contains dilute sulphuric acid into which dips a rod of amalgamated zinc.

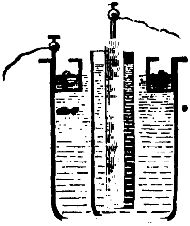

FIG. 98.—Daniell's Cell.

The chemical action, on which the production of the electric current depends, begins in the neighbourhood of the zinc rod, and is of precisely the same nature as that which has been seen to occur in a simple voltaic cell. There is a simultaneous production of zinc sulphate and free hydrogen. This hydrogen similarly decomposing the neighbouring molecule of sulphuric acid results in the liberation of more hydrogen, this action being reproduced throughout the row of sulphuric acid molecules between the zinc rod and the porous vessel. The free hydrogen which is there formed, and which would in the simple voltaic cell be deposited on the copper plate, passes through the pores of the earthenware partition and comes into contact with the copper sulphate solution. A chemical reaction ensues. The hydrogen, being just liberated from combination, is in a particularly active state, known as the *nascent* condition, and is under these circumstances able to effect the decomposition of the copper sulphate. We can represent the change which follows thus :—

The free HYDROGEN acts upon the COPPER SULPHATE decomposing it to form

SULPHURIC ACID and COPPER.

The same reaction is more shortly expressed by a chemical equation :—

$$H_2 + CuSO_4 = H_2SO_4 + Cu.$$

The free copper must now be supposed to act upon the neighbouring molecule of copper sulphate decomposing it, forming copper sulphate, which contains the same proportions of the constituent elements of course, but not the identical atoms. Just as in the case of the sulphuric acid in the inner cell, so here in the outer, there is a series of decompositions and recompositions, with the final result that there is eventually a deposit of free copper formed upon the inside of the copper vessel, but no hydrogen bubbles.

The cell continues to develop an electric current at the expense of the zinc, which is gradually dissolved. Throughout the passage of the electric current round the wire joining the zinc rod and the copper vessel, there is a continuous deposition of metallic copper on the inside of the copper vessel, which consequently gets heavier and heavier.

This form of cell does not give a very strong current, but the great advantage about it is that the strength of the current is very uniform. It is largely used by the Post Office authorities for telegraphic purposes.

2. **Bunsen's and Grove's Cells.**– *The hydrogen is got rid of by oxidation.* The only difference between these two kinds of voltaic cells is that, whereas, in the former a piece of gas carbon replaces the copper plate, in the latter there is a plate of platinum. Owing to the cheapness of gas carbon Bunsen's cell is the most commonly used.

In Bunsen's cell there are two earthenware vessels, the inner smaller one is porous and is filled with strong nitric acid, into which the piece of gas coke C dips. The outer vessel contains dilute sulphuric acid, and in it is placed the zinc plate Z, which is usually made cylindrical in shape. The arrangement of the parts is easily understood by a reference to Fig 99.

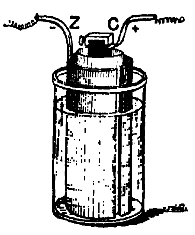

FIG. 99.—Bunsen's Cell.

The chemical action begins at the zinc plate and the result of it is that, as in Daniell's cell, free hydrogen passes through the porous vessel and comes into contact with the strong nitric acid contained in it. Nitric acid is remarkable for the large quantity of oxygen it contains and for the ease with which it parts with it. It is hence known as an *oxidising* *agent.* Consequently, when the free hydrogen comes into contact with the nitric acid the chemical change which occurs is that the nitric acid gives up some of its oxygen, and this combines with the free hydrogen to form water ; or, as it is commonly expressed, the hydrogen is *oxidised* to water by the oxygen of the nitric acid, which is itself *reduced*, first to nitrous acid and eventually to the lower oxides of nitrogen. Thus is the free hydrogen got rid of.

Leclanché's Cell.—This cell (Fig. 100), which we shall describe because of its familiar use in houses for working electric bells, is different in two or three respects from those which have already been dealt with. There is no acid used in it. The liquid which acts upon the zinc plate Z is a solution in water of a salt called ammonium chloride or sal-ammoniac. The result of the chemical changes which take place when the liquid acts upon the zinc is that a compound known as a double chloride of zinc and ammonium is formed, and two gases, ammonia and hydrogen, are evolved. To get rid of the hydrogen the positive pole C of the cell, which is made of carbon, is surrounded with black oxide of manganese, a compound rich in oxygen. The carbon plate and the fragments surrounding it are enclosed in a porous pot P, like that of the Daniell's cell. The hydrogen which it is necessary to get rid of combines with some of the oxygen of the manganese compound and is changed into water, while the manganese compound is gradually reduced.

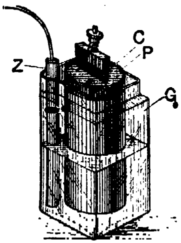

FIG. 100.—Leclanché's Cell.

Positive and Negative Poles.—In each one of the cells which has been described the chemical action in the cell has started at the zinc plate, and has proceeded from molecule to molecule throughout the liquid until the copper plate, or its equivalent, has been reached. The plate where the chemical action begins, and the electric current starts *in* the liquid, is called the **negative (–) pole**. The plate where the chemical action in the cell is completed and where the current starts *outside* the liquid is called the **positive (+) pole**. These terms are easily understood by reference to Fig. 94, which represents a **closed circuit**, as it is called. If the wires from the two plates are unconnected the current is said to be **open,** and no electric current circulates.

Work done by the Current.—The only work which the electric circuit performs in a closed circuit, such as is shown in Fig. 94, is of a useless kind, for it is completely used up in warming the connecting wire and the liquid.

> Expt. 187.—Connect the wires from the carbon and zinc poles of a Bunsen's battery to the ends of a thin piece of platinum wire. Notice that when the circuit is closed the platinum wire becomes red hot.

But there are many kinds of useful work which such a current can perform.

> Expt. 188.—Fit up a Bunsen's cell, taking care that there is a bright metallic connection at each binding screw. Join the wires from the carbon and zinc poles respectively to the two binding screws of an electric bell. Notice the ringing.
>
> Expt. 189.—Similarly attach the wires to the binding screws of an electro-magnet. Notice (*a*) its lifting power, by attaching gradually increasing weights to the hook of the armature; (*b*) remove the armature, support the electro-magnet vertically, with the poles or ends of the magnet upwards, and place a sheet of cardboard, or a plate of glass, on them. Sprinkle fine iron filings evenly over the cardboard or glass, gently tap the latter, and notice their arrangement in lines, which proceed radially from the poles, and which are called *lines of force* (Fig. 102).

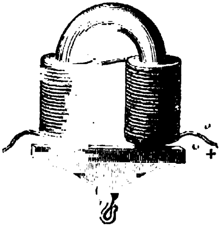

Fig. 101.—An Electro-Magnet with Armature.

EXPT. 190.—Dip the free ends of the wires into some very dilute sulphuric acid, contained in a beaker, and notice that bubbles of gas are formed on each wire.

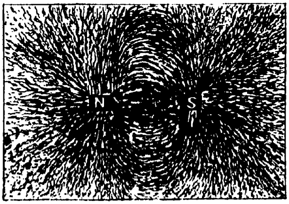

FIG. 102.—Lines of Force shown by arrangement of Iron Filings upon a Glass Plate lying upon the poles of an Electro-Magnet.

CHIEF POINTS OF CHAPTER XII.

Fundamental Facts.—(1) Commercial zinc is acted upon by dilute sulphuric acid with evolution of hydrogen. (2) There is no such action when pure zinc or amalgamated zinc is used. (3) If strips of pure or amalgamated zinc and copper be placed in dilute sulphuric acid and the metals connected outside the liquid by a copper wire, the zinc is acted upon and hydrogen bubbles appear on the copper strip.

Properties of the Connecting Copper Wire.—(1) The wire deflects a magnetic needle placed under it. (2) The wire attracts iron filings. (3) After a time the magnetic power of the wire becomes feebler. (4) If the wires from the copper and zinc plates be dipped into acidulated water, instead of being joined, bubbles of gas appear on the ends of each wire.

Potential.—The electric current passes along the wire because of the *difference of potential* between its ends. The end of the wire in connection with the copper plate is at a *higher potential* than that joined to the zinc plate. Potential is analogous to water-level and temperature.

Electro-motive Force.—Since electric currents flow from places of high potential to places of low potential, and the strength of the current is directly proportional to the *difference of potential* between the places, it is customary to speak of this difference of potential as *electro-motive force.* Electro-motive force can be directly measured with a Quadrant Electrometer.

Simple Voltaic Cell.—The arrangement of a plate of pure or amalgamated zinc and one of copper in dilute sulphuric acid constitutes a simple voltaic cell. The hydrogen which collects on the copper plate soon stops the flow of the current. This effect is called polarisation.

Voltaic Cells in which Polarisation is prevented.—(*a*) *By mechanical means.* **Smee's Cell.**—A platinised silver plate is substituted for the copper one. Owing to the roughness of such a plate the hydrogen bubbles stream off it.

(*b*) *By chemical means.* **Daniell's Cell.**—The hydrogen, instead of being allowed to collect, is made to act upon a solution of copper sulphate which it decomposes, forming sulphuric acid with it and liberating metallic copper. This copper being deposited on the copper plate does not interfere with the passage of the current.

Bunsen's and Grove's Cells.—Here the hydrogen is got rid of by oxidation. This is effected by causing it to come in contact with strong nitric acid, which being rich in oxygen, changes the hydrogen into water, or *oxidises* it. The nitric acid is itself *reduced*, first to nitrous acid and eventually to oxides of nitrogen. In Bunsen's cell a block of gas carbon takes the place of the copper plate of the simple voltaic cell, and in Grove's a plate of platinum.

Leclanché's Cell is the one commonly used in houses for ringing electric bells. The liquid used is a solution of ammonium chloride. The hydrogen formed by the action of the liquid on the zinc plate is got rid of by causing it to combine with oxygen from black oxide of manganese which surrounds the carbon pole.

Positive and Negative Poles. The *negative* (−) *pole* is usually of zinc and is the one where chemical action begins and consequently where the electric current starts *in* the liquid.

The *positive* (+) *pole* is the plate where the chemical action is completed and where the current starts *outside* the liquid.

QUESTIONS ON CHAPTER XII.

1. What happens under the following circumstances :—
   (*a*) A piece of commercial zinc is plunged into dilute sulphuric acid.
   (*b*) Pure zinc is treated as in (*a*).
   (*c*) Pieces of amalgamated zinc and ordinary copper are placed parallel to one another in dilute acid and connected by a copper wire outside the liquid.

2. Before fitting up a simple voltaic cell both plates are accurately weighed. They are weighed again after the current has passed for ten minutes. What differences will be noticed and how would you account for them?

3. What properties do you associate with the wire connecting the plates of a simple voltaic cell?

4. Explain as fully as you can what is meant by potential. State by the use of this term how it is that an electric current passes round a closed circuit.

5. It is noticed that after a short time the strength of the electric current from a simple voltaic cell becomes very much diminished. How do you account for this?

6. Explain what is meant by the *polarisation of a cell.* What devices are employed to get rid of polarisation?

7. Give a description of the voltaic cells named after the following physicists :—

(*a*) Daniell.
(*b*) Grove.
(*c*) Bunsen.
(*d*) Leclanché.

8. What chemical changes occur in a simple voltaic cell?

9. What is the object of the nitric acid in a Grove's cell?

10. Sketch the chemical changes which occur between the zinc and copper plates in a Daniell's cell.

11. Compare the flow of an electric current with the passage of heat from a hot to a cold body.

12. Describe experiments to show that the electric current can do work.

CHAPTER XIII

CHEMICAL CHANGE INDUCED BY THE ELECTRIC CURRENT

How the Passage of an Electric Current can be recognised. The Galvanoscope.—Before proceeding to study the chemical changes induced by an electric current it is very desirable that a method should be learnt by means of which the passage of a current of electricity through a circuit can easily be perceived. Expt. 191 provides such a means, and if it be modified somewhat we are in possession of a sure plan of recognising even small currents.

EXPT. 191.—Instead of a single wire such as was used in Expt. 184, substitute an arrangement like that shown in Fig. 103, where a light magnetic needle hangs inside a coil of wire. The effect of coiling the wire is, roughly, to multiply the effect of the single wire by the number of turns there is in the coil. Connect the ends of the coil with a copper and an amalgamated zinc plate and dip the plates into dilute sulphuric acid. Compare the effect on the needle in this case with that obtained by passing the current along a single wire.

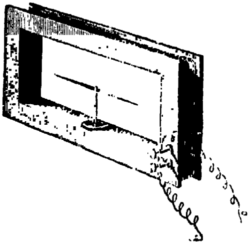

FIG. 103.—A Galvanoscope.

EXPT. 192.—Substitute an **astatic pair** for the single magnetic needle in the last experiment. The astatic pair consists of two equally strong magnets arranged parallel to one another, with opposite poles in the same direction, and permanently joined together by a suitable connection such as copper wire. Notice that in whatever direction the astatic pair points there is no disposition for it to change this direction; whereas the single magnetic needle will only rest in one direction, called the magnetic meridian. Of course, if one of the needles is magnetically stronger than the other they will not exactly counterbalance one another.

EXPT. 193.—Arrange the astatic pair and a coil of wire as in Fig. 105. Notice that a very slight current, such as that obtained by dipping a brass pin and a steel needle into dilute sulphuric acid, is enough to cause a deflection of the needle.

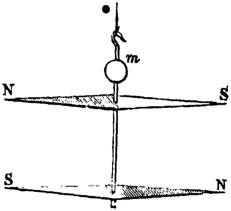

FIG. 104.—An Astatic Pair of Magnetic Needles.

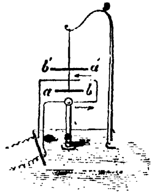

FIG. 105.—An Astatic Galvanoscope.

The Passage of the Electric Current through Liquids.—EXPT. 194.—Fit up a Bunsen's cell for the generation of an electric current. Attach pieces of platinum foil, by means of suitable binding screws, to the ends of two copper wires. Attach one of these wires to one pole of the battery. Connect the other pole to one of the binding screws of a simple galvanoscope, and to the other screw of the galvanoscope attach the remaining wire with the platinum plate on the end (*see* Fig. 107). Dip the platinum plates—1st, into some mercury, and notice there is a great deflection of the needle of the galvanoscope and no alteration of the mercury; 2nd, into some turpentine, and notice there is no deflection of the needle; 3rd, into some acidulated water, and notice there is a smaller deflection than in the first case, and at the same time there are bubbles of gas given off from both platinum plates.

1st Case. Passage of the Current through Mercury.—The student will shortly learn (Chap. XIV) that mercury is one of the chemical elements, and that it is so regarded because it can neither be decomposed by the passage of an electric current nor by any other known methods. The great deflection of the needle of the galvanoscope reveals the fact that a considerable current passes through its coil of wire, hence we say that mercury is a good conductor of the electric current, or expressing the same truth in other words, that it offers very little **resistance** to the flow of the current.

Similarly, it would be found that other metals, when in the liquid condition, which they can assume if the temperature is raised sufficiently, are also good conductors of the electric current.

2nd Case. Passage of the Current through Turpentine.—There is in this case no deflection of the needle of the galvanoscope; it is therefore evident that no current passes through the coil of wire round the needle, and since the battery is arranged precisely as in the previous experiment with the mercury, the explanation must be that the turpentine prevents the flow of the electric current round the circuit. Turpentine is consequently known as a **non-conductor**, a class of bodies which also includes such liquids as petroleum and other oils.

3rd Case. Passage of the Current through Acidulated Water.—Here we have the current conducted and the liquid decomposed by the passage of the current. This is the condition of things in all liquid compounds which conduct the electric current. Such a decomposition as this is known as **electrolysis**, and we shall have to study this case more fully.

Electrolysis of Water.—Pure water is a very bad conductor of the electric current, and hence, as in Expt. 245, it is necessary to add a drop or two of acid to make it conduct. To understand exactly the result of the passage of the electric current, some means must be devised by which the gases which appear at the platinum plates can be collected. Such an arrangement constitutes what is known as a **voltameter.** [This piece of apparatus will be again referred to later on (Chap. XVI).] A convenient pattern to use for the decomposition of water consists (Fig. 106) of a glass vessel in the bottom of which are fixed two slips of platinum connected, by means of copper wires, with two binding screws. Before connecting the binding screws with the poles of a battery, acidulated water is poured into the vessel and two glass tubes, of exactly equal size and carefully divided into equal volumes by divisions etched on the glass, are completely filled with acidulated water and inverted over the platinum plates, as shown in the illustration. When the binding screws are connected with the poles of a battery, consisting of two or three Bunsen's cells, bubbles of gas are immediately seen to appear on the platinum plates, and in a few minutes it will be observed that a quantity of gas has collected in each tube. If we allow the electric current to pass round the circuit for 15 or 20 minutes and then measure the volume of the gas which has collected in each tube, it will be found that the tube over the platinum plate which is connected with the zinc pole of the battery contains twice as much gas as that over the plate connected with the carbon pole. Moreover, if the bottom of the tube containing the larger amount of gas be covered with the thumb and the tube be lifted out of the liquid, inverted, and a light applied to the gas, it will be found to burn, showing it to be hydrogen (Chap. XVI). Similarly, if a glowing splinter of wood be plunged into the other gas it will be rekindled, showing it to be oxygen. Hence we see that the passage of a sufficiently strong electric current through water causes it to be decomposed into hydrogen and oxygen, and that twice as much hydrogen by volume as oxygen is formed. Or,

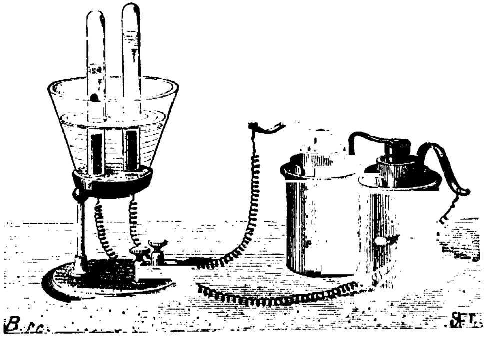

FIG. 106.—Electrolysis of Water.

Water is decomposed by the electric current into Hydrogen [2 vols.] and Oxygen [1 vol.]

fact which can be at once simply expressed by chemical ymbols :—

$$H_2O = H_2 + O.$$

Terms used in describing Electrolysis.—It is customary in peaking of the decompositions effected by the electric current to se certain terms originally adopted by Faraday. The liquid vhich conducts the electric current, and is itself decomposed, is nown as the **electrolyte** ; the platinum plates in the voltameter lescribed, or, generally speaking, the ends of the wires coming rom the poles of the battery, are called **electrodes**. Names are iven to each of these to distinguish them. That by means of vhich the current *enters* the electrolyte, or what is the same hing, the electrode in connection with the carbon pole of a Bunsen's battery, or the positive (+) pole of any battery, is alled the **anode**. The electrode by means of which the electric urrent *leaves* the electrolyte or that in connection with the inc or negative (−) pole of the battery, is called the **kathode**. he atoms into which the molecules of the electrolyte are lecomposed are referred to as **ions**. Evidently the ions are nvisible during their passage through the electrolyte. Those ons which collect at the anode are spoken of as **anions**, those ollecting at the kathode are called **kathions**. It will be clear o the reader that the direction in which the kathions move hrough the liquid is that which we have spoken of as the lirection of the current, namely, from the anode to the kathode Fig. 107).

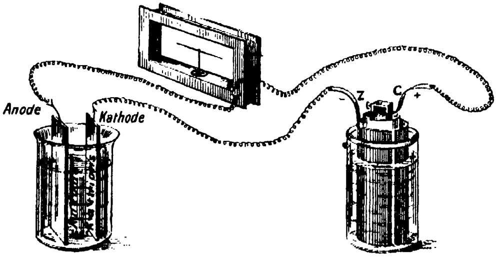

FIG. 107.—Current passing through Galvanoscope and an Electrolyte.

Electrolysis of Copper Sulphate.—EXPT. 195.—Make a strong solution of copper sulphate (blue vitriol) in water. Pour some into a beaker. Dip two platinum plates connected by copper wires, as previously described (Expt. 194), into the solution and notice that after a few minutes a deposit of copper is found on the kathode, and that bubbles of gas (which if collected and tested are found to be oxygen) are seen to rise from the anode.

The passage of the electric current through copper sulphate is most easily expressed by means of chemical symbols, and we shall consequently suppose the reader has studied the section dealing with such symbols before attempting to understand this description. Copper sulphate is constituted as shown in the formula $CuSO_4$. The effect of the passage of the electric current is to cause it to decompose according to the equation into copper and the group of elements, SO_4, which is, however, immediately acted upon by the water in which the blue vitriol is dissolved as shown in the second equation :—

$$CuSO_4 = Cu + SO_4$$
$$SO_4 + H_2O = H_2SO_4 + O.$$

Hence we see that there is a continual withdrawal of metallic copper from the solution by the passage of the electric current, and the liquid is gradually converted into sulphuric acid. This can be demonstrated by its effect on a piece of blue litmus paper.

If instead of having two platinum plates as electrodes we substitute copper plates, we have a slightly different state of affairs.

EXPT. 196.—Arrange the apparatus as described in Experiment 195, substituting copper plates for the platinum ones there used, and weigh the copper electrodes before passing the current. After the current has passed for, say, ten minutes, break the circuit and weigh the electrodes again. Notice the anode has lost in weight by a certain amount, and that the kathode has increased in weight *by the same amount*. There is no evolution of oxygen.

Secondary Results.—The evolution of oxygen which is noticed in Expt. 195 is a result of what is known as a **secondary** kind. The liberated group of elements, SO_4, is the first effect of the passage of the electric current, but, as we have seen, this immediately reacts with the water present and oxygen appears at the anode. In the next illustration of the chemical action of the electric current we have two secondary products formed.

Electrolysis of Sodium Sulphate.—Expt. 197.—Make a strong solution of sodium sulphate (Glauber's salt), Na_2SO_4, in water. Place a porous pot out of a Bunsen's cell into a large beaker. Pour some of the sodium sulphate solution into the porous vessel and also fill the beaker to the same height with the solution. Colour these portions of the solution with litmus, and then redden the solution in the porous vessel by the addition of a drop or two of acid. Plunge platinum electrodes into the solution; one in the porous vessel and the other in the beaker. Connect that in the former with the zinc or − pole of the battery and that in the beaker with the carbon or + pole. After a time it will be noticed that the solution in the porous vessel becomes blue, while that in the beaker turns red. It will also be noticed that gases are given off at each electrode; that from the one in the porous vessel will, if collected and examined, be found to be hydrogen, that from the other electrode oxygen.

Explanation of Electrolysis of Sodium Sulphate.—The experiment really consists of three chemical reactions. First the sodium sulphate is decomposed, as the following equation shows:—

$$Na_2SO_4 = Na_2 + SO_4.$$

But at the moment of their liberation the two groups, Na_2 and SO_4, are acted upon by the water in which the salt is dissolved. As the student will learn, sodium acts upon water, forming sodium hydrate and liberating hydrogen, a change which is most simply expressed by an equation:—

$$Na_2 + 2H_2O = \underset{\substack{\text{sodium} \\ \text{hydrate.}}}{2NaHO} + H_2$$

It is this hydrogen, which is the result of a secondary reaction, which is liberated at the kathode. The sodium hydrate, being alkaline, has the power of turning reddened litmus blue again, a phenomenon which was observed in performing the last experiment.

The third reaction which takes place is the action of water upon the liberated group, SO_4, which has been described in the electrolysis of copper sulphate. It accounts for the reddening of the solution in the beaker and for the evolution of oxygen which takes place at the anode.

CHIEF POINTS OF CHAPTER XIII.

The **Galvanoscope** is a convenient apparatus for recognising the passage of an electric current.

The Passage of the Electric Current through Liquids.—

(*a*) Liquid metals conduct the current without being decomposed.

(*b*) Certain liquids, like turpentine, will not conduct the electric current and are consequently not decomposed by it.

(*c*) Compound liquids, which, like acidulated water, conduct the electric current, are decomposed by its passage.

Electrolysis is the term used to refer to condition (*c*) above. It means the process by which electric currents pass through compound liquids and so cause them to be decomposed. The liquid which conducts the electric current and is itself decomposed is known as the **electrolyte.**

The ends of the wires coming from the poles of the battery are called **electrodes**, that by which the current enters the electrolyte is known as the **anode**, that by which it leaves the **kathode** (or cathode).

The atoms into which the electrolyte is decomposed are called **ions.**

The ions which collect at the anode are the **anions** ; those which collect at the kathode the **kathions** (or cathions).

Electrolysis of Water, H_2O—

WATER is decomposed by the electric current into HYDROGEN [2 vols.] and OXYGEN [1 vol.]

a fact which can be at once simply expressed by chemical symbols :—

$$H_2O = H_2 + O.$$

Electrolysis of Copper Sulphate, $CuSO_4$.—First, the copper sulphate is broken up into copper and the group of elements SO_4, thus :—

$$CuSO_4 = Cu + SO_4$$

Then a *secondary reaction* takes place between the group of elements SO_4, and the water in which the copper sulphate is dissolved :—

$$SO_4 + H_2O = H_2SO_4 + O.$$

The final result is that copper is deposited at the kathode and bubbles of oxygen are evolved at the anode.

Electrolysis of Sodium Sulphate, Na_2SO_4.—First, the sodium sulphate is decomposed into sodium and the group of elements SO_4, thus :—

$$Na_2SO_4 = Na_2 + SO_4.$$

This decomposition is followed by two secondary reactions : the sodium at once acts upon the water present forming sodium hydrate and evolving hydrogen :—

$$Na_2 + 2H_2O = 2NaOH + H_2 ;$$

and the group of elements, SO_4, reacts with water as described in the case of the electrolysis of copper sulphate. The final result is that hydrogen is evolved at the kathode and oxygen at the anode.

QUESTIONS ON CHAPTER XIII.

1. Describe fully what takes place when each of the following liquids is included in a circuit round which a strong electric current is circulating :—
   (*a*) Liquid mercury.
   (*b*) Petroleum.
   (*c*) Acidulated water.
2. What means would you adopt to recognise the passage of an electric current through a wire ?
3. What do you understand by *electrolysis ?* Give an account of the electrolysis of water.
4. Explain the terms :—electrolyte, anode, kathode, ion.
5. What happens precisely when an electric current is passed through a solution of copper sulphate in water. What appears at the anode, and how do you account for its appearance ?
6. Explain by the aid of chemical symbols the various chemical changes which take place during the passage of the electric current in the last question.
7. Write a short essay on the electrolysis of a solution of sodium sulphate.
8. Clearly explain the appearance of hydrogen at the kathode when an electric current is passed through a solution of Glauber's salt.
9. After the passage of an electric current through a solution of sodium sulphate, the liquid round the anode is found to turn a blue litmus paper red. Explain this.

CHAPTER XIV

INTRODUCTORY.—SIMPLE CHEMICAL OPERATIONS

Physical and Chemical Changes.—Matter is subject to two kinds of change. Hitherto we have been chiefly concerned with those which influence the properties of matter, leaving its composition unaltered. It has been seen that a body, such as a piece of iron, may gradually increase in temperature, changing from cold iron to hot, and, becoming hotter and hotter, may change in colour, passing from a dull gray to red, and from red to almost white, becoming incandescent and emitting light rays. But if left to itself the iron will begin to cool, passing through the same changes in the reverse order until it reassumes precisely the former condition; and in all these changes the weight of the iron remains unaltered. Or, again, we might take a piece of soft iron, and, having wound silk-covered copper wire round it several times, pass an electric current through the wire. It would be found, on examining the iron, that new properties had been imparted to it, that it was now able to pick up other pieces of iron, or had become magnetised. If the electric current be discontinued, the new power, too, disappears. Such changes as these, where the substance or composition of the body remains unchanged, are known as *physical changes.* On the other hand, if a piece of iron be left exposed to damp air for some hours it becomes covered with a reddish-brown powder, which the most superficial examination will show is a different substance from the iron with which we started. There are a very large number of changes of the same kind as this continually taking place around us. When gunpowder explodes, we have an abundance of smoke formed and a black residue left behind, and it is easy to see that the smoke and deposit are quite unlike the gunpowder before the flash. Such changes as these are called *chemical changes*. It is with changes of this second kind that Chemistry is concerned, and we may define this science thus: *Chemistry is that branch of knowledge which deals with chemical changes; those, namely, which result in the formation of new substances with new properties.*

Chemical Elements.—The result of a large number of experiments made from time to time by different chemists has been to show that there are upwards of seventy different forms of matter which can by no known methods be broken up into anything simpler. By this is meant that if any one of these be selected and treated in any way with which chemists are familiar —for example, if it were subjected to a very high temperature— we should find it impossible to get anything having properties different from those of the substance itself; bodies of this simple kind are called **elements**.

But it must be carefully borne in mind that, as the methods which chemists adopt become more and more refined, it is quite likely that some of these may be found to be wrongly regarded as elements. Up to the time of Davy (1807) the substances soda, potash, and lime were regarded as elements. He found, however, that they could be split up into simpler constituents. From soda he obtained a soft metal, sodium, and two gases, oxygen and hydrogen, and from that time, of course, soda could not be regarded as an element. Similarly, if at any future time it should be found that any of the forms of matter which we call elements can be split up into simpler bodies with different properties, the element which is thus decomposed will have to be struck off the list.

A list of the substances now regarded as elements is given at the end of Chap. XIX.

Metals and Non-Metals.—A good many of these substances are possessed of certain distinctive characters in which they resemble one another. They have a bright lustre, a high specific gravity (*see* p. 96), are good conductors of heat and electricity, and are known to chemists as **metals**. There is no difficulty in deciding in a large number of instances that the substances possess the characters of a metal, and the student will immediately think of gold, silver, copper, iron, &c. Other bodies, however, are quite as plainly not of this class; they have no lustre, they are not heavy, nor do they conduct heat and electricity well. These are spoken of as **non-metals**, and phosphorus, sulphur, and carbon will serve as good instances. But the line between the two classes is not a hard and fast one, for one or two substances possess some of the properties which distinguish a metal, and yet for other reasons, which the student will understand better later, are not classed with the metals, but with the non-metals. Arsenic may be cited as an instance of a substance which possesses properties common to both classes.

Chemical Operations. A number of chemical operations are so frequently employed, that it will be well at once to familiarise ourselves with them, and we may do this by the following experiments.

EXPT. 198.—Place a piece of sugar in water, and note that it soon disappears, and has given a sweet taste to the whole of the water, so that the particles of the sugar must be spread through the entire mass of the water.

The sugar is said to have *dissolved* and formed a *solution* of sugar. Similarly, a very great number of substances can dissolve, or are soluble, in water, but not all to the same extent, and when no more of the substance can be dissolved the solution is said to be *saturated*.

EXPT. 199.—Weigh out 50 grams of each of the following substances:—finely powdered nitre, sugar, salt, and to each add water, in small quantities, with vigorous shaking after each addition. Determine thus the quantity of water necessary to form a saturated solution of each.

Solution is a Physical Change.—EXPT. 200.—Weigh out a quantity of salt in an evaporating basin and dissolve it in water. Heat gently over a Bunsen burner so that the water boils and evaporates away completely. Note that a white solid remains in the basin, and again weigh. Satisfy yourself that the weight is equal to the weight of the basin and salt before solution, and that the solid left is still salt.

Hence during solution[1] we have in general no *chemical change*, but merely a change of physical state, although we shall find later that in certain cases solution may be accompanied by chemical changes.

By very gently warming water or other liquid, or even by allowing it to remain exposed to the air for some time, the liquid passes off as vapour without actually boiling, as the student may readily see by a simple experiment, and we speak of this as **evaporation.**

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[1] The *act* of solution is now commonly spoken of as dissolution, the term solution being restricted to the product formed by the dissolution.

Distillation.—We also see that we can by the evaporation of the solvent, *i.e.* the liquid in which the solution takes place, separate it from the dissolved substance, and this is very frequently used not only for obtaining the dissolved substance but also for the purification of liquid from dissolved solid material, since if the steam which is formed by boiling water containing any of these dissolved substances be condensed, the water formed is quite pure. To obtain pure water from any kind of water, then, whether fresh or salt, all that has to be done is to boil it and condense the steam which is given off. The dissolved materials are all left behind in the vessel in which the boiling takes place. An arrangement for condensing steam or vapour is shown in Fig. 108. The steam that is driven off from the water in the flask passes through a long tube kept cool by being surrounded with water, and is thus condensed.

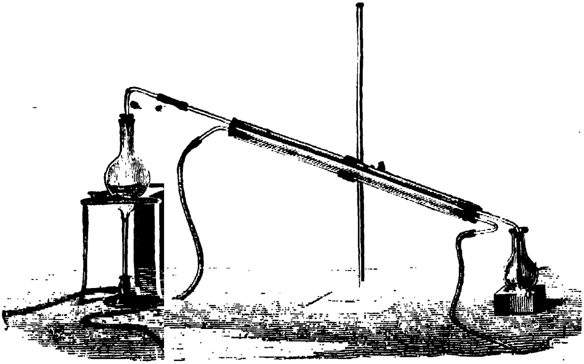


FIG. 108.—The Distillation of Water.

EXPT. 201.—Evaporate a little distilled water in a platinum or porcelain crucible. Notice the absence of any residue. Repeat the experiment with tap-water, and note the residue.

Separation of Solids.—Filtration.—We may also use the solvent power of water to separate solids which are insoluble from those which are soluble.

EXPT. 202.—Mix well together some fine sand and salt; place the mixture in a glass flask with sufficient water to dissolve the salt, and shake vigorously. Take a circular piece of filter-paper and fold it in two, in the form of a semicircle, then again to the form of a quadrant; open out into the form of a hollow cone and fit it into a glass funnel (Fig. 109). Pour the liquid through this filter and note that the solution of salt runs through perfectly clear leaving the insoluble sand in the filter paper. Evaporate off the water from the salt solution and obtain the salt.

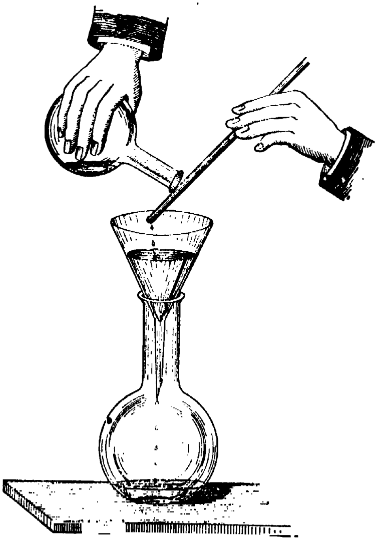


FIG. 109.—Filtration.

The above method of *filtration* is that which we usually adopt in chemical operations when we wish to separate insoluble suspended material from a liquid, either for the isolation of the solid or the purification of the liquid.

Solvent Power of Hot and Cold Water.—We may now find whether temperature has any influence on the solvent power of water, that is, does hot water dissolve more or less of a solid than cold water?

> EXPT. 203.—Place a quantity of powdered nitre in water and allow it to stand for some time with frequent vigorous shaking, so that a cold saturated solution is formed. Now heat the solution and see whether more solid dissolves or not.

Do the same experiment with other solids and satisfy yourself of the truth of the following statement, that in general hot water dissolves more of a solid than cold water, or in other words, that the solvent power of a liquid increases with rise in temperature.

> **Crystallisation.**—EXPT. 204.—Allow a hot saturated, or nearly saturated, solution of nitre obtained as above to cool slowly, and observe that the nitre separates out from the solution as clear glassy solids, which may be observed to increase in size as the solution cools. Examine a few of these and see that they are bounded by plane surfaces.

These are known as *crystals* and their formation as *crystallisation*. Crystallisation usually occurs when a dissolved solid separates out from its solution, either by cooling of the solution, or by the evaporation of the solvent; and the slower the evaporation, or more gradual the cooling, the larger and more perfect are the crystals obtained. This may be easily verified by rapidly cooling, in a stream of cold water, a hot saturated solution of nitre. We shall also find later that crystals may also be formed during the solidification of a melted solid, especially when it is allowed to cool and solidify slowly and undisturbed.

> **Precipitation.**—EXPT. 205.—Dissolve some lead acetate in distilled water; add to it a solution of salt. Observe the formation of a white powder in the previously clear liquid.

Such precipitation always happens if to a solution we add some material capable of converting the dissolved substance into a new product which is insoluble in water, and which therefore separates out as a powder. This insoluble powder is called a *precipitate*, and is said to be *precipitated*.

Solution of Liquids.—EXPT. 206.—Shake up a little water and ether. Observe the separation into two layers; ether above and water below. Pour the mixture into a funnel the bottom of which is closed by your finger. Allow a little of the lower liquid to run into a test-tube. Insert into the test-tube a cork fitted with a glass tube. Boil the liquid and apply a light to the tube, and observe that an inflammable gas is at first evolved, but after a while only steam passes off.

The lower layer of water hence contained a quantity of dissolved ether, while it may be also proved that the upper layer of ether contained dissolved water.

EXPT. 207.—Add alcohol to water, and shake. Observe that they do not separate into two layers, but form one liquid.

When liquids, as in the case of alcohol and water, form one homogeneous liquid, they may be said to be mutually soluble in all proportions. In such a case the liquids are said to be *miscible* or to *mix* with one another.

Hence ether and water are mutually soluble but do *not* mix.

Mixtures and Compounds.—Substances which contain more than one element may be either *compounds* or *mixtures*, and the difference between these will be best understood by performing the following experiment:—

EXPT. 208.—Mix together some copper filings and powdered sulphur, and examine the result, which is a *mixture* of copper and sulphur. See that its colour lies between the yellow of sulphur and the red of copper, and that the particles are quite distinct. Wash a little in a gentle stream of water, and observe that the sulphur gets washed away more readily than the copper, leaving the latter as a residue. Shake up with a little of the liquid known as carbon disulphide, and see that the sulphur dissolves but the copper is left; pour off the solution and allow it to evaporate, and observe that the sulphur is left.

Now again mix the two ingredients in a porcelain crucible, using a weighed quantity of copper, and heat either in the open air or in a draught cupboard. Observe that the sulphur melts and some burns away. Add more sulphur, and again heat until no more sulphur burns away, and do this three or or four times. Weigh the product so obtained, and then examine it carefully. Try to separate the sulphur and copper by washing, and try to dissolve the sulphur out by carbon disulphide. You will now find that you cannot separate the two components, neither can you distinguish the individual particles.

The substance so formed is a *compound* of copper and sulphur, so that during the heating a change has taken place, with the formation of a new substance with properties of its own. This is an example of a chemical change - the result of chemical action. The new substance is a chemical compound. If you have taken the weight carefully you will have also found that 100 parts of copper form, roughly, 125 parts of the compound, that is, unite with 25 of sulphur, and the result does not depend on the actual weight of copper and sulphur used, all excess of sulphur over this being always burnt away. We thus find the following differences between a mixture and a compound :—

MIXTURE.—In a mixture the components exist side by side and can be separated by simple mechanical methods. The ingredients may be present in any proportions, and the properties of the mixture are intermediate between those of the constituents.

COMPOUND.—In a compound the components cannot be separated by the simple means available in the case of mixtures. The properties of the compound are quite different from those of the constituents, and these constituents are always present in certain definite proportions which for each compound are invariable.

In all cases of chemical action it is most important to remember that the total weight remains absolutely unchanged, that is, the total weight of *all* the products is exactly equal to the total weight of all the components forming these products.

CHIEF POINTS OF CHAPTER XIV.

Physical Changes are those in which the composition of the body experiencing the change remains unaltered. The science concerned with these changes is called **Physics**.

Chemical Changes are those which result in the formation of new substances with new properties. The study of such changes is called **Chemistry**.

Chemical Elements are kinds of matter which can, by no known means, be broken up into anything simpler. They can be subdivided into **metals** and **non-metals.**

Chemical Operations.—**Solution** is the process by which some substances, when placed in water or other liquids, disappear and their particles spread through the entire mass of the water or other liquid.

When no more of the substance will *dissolve* the liquid is said to be *saturated.* Solution is generally a physical change and is unaccompanied by any changes of weight.

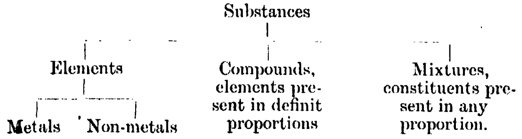
Distillation is the process by means of which we are able to separate the dissolved substance from the solution containing it. The liquid itself is recovered by condensing the vapour.

Filtration consists in utilising the porosity of unglazed paper or other material for separating insoluble substances from solutions with which they are mixed; or, it is a means of purifying liquids and isolating solids.

Crystallisation denotes the separation of a dissolved solid in masses having regular shapes from a solution as the liquid evaporates or cools. The regular solids formed are called crystals.

Precipitation means the formation and separation of an insoluble substance from a solution.

Compounds and Mixtures.—


Substances
Elements
Metals
Non-metals
Compounds, elements present in definit proportions
Mixtures, constituents present in any proportion.

In a **mixture** the components exist side by side and can be separated by simple mechanical methods. The ingredients may be present in any proportion and the properties of the mixture are intermediate between those of the constituents.

In a **compound** the components cannot be separated by the simple means available in the case of mixtures. The properties of the compound are quite different from those of the constituents and those constituents are always present in certain definite proportions which for each compound are invariable.

QUESTIONS ON CHAPTER XIV.

1. How may muddy water be (*a*) made clear, (*b*) purified from dissolved material?

2. Explain the differences between a mixture and a compound.

3. What do you mean by a "physical" and what by a "chemical" change? If a substance changes under the influence of heat how would you endeavour to find out whether the change is chemical or merely physical?

4. A mixture of salt and powdered glass is given to you. How could you (*a*) obtain separately the two constituents, (*b*) find the

quantity of each in the mixture? Could you separate sugar from salt in the same way?

5. What do you mean by a saturated solution? How would you prepare a saturated solution at a given temperature? What is generally the effect of cooling a saturated solution?

6. A white powder is shaken up with water. How would you ascertain whether any of it dissolves?

7. What is meant by the statement that two liquids mix? Give examples. Can a liquid be soluble in another liquid without being capable of mixing? If so, give an example.

8. How may fresh water suitable for drinking be obtained from sea water?

9. By what means would you endeavour to find the quantity of a given solid which a given quantity, say 1 gallon, of water is capable of dissolving?

10. What is meant by "precipitation?" Under what circumstance may precipitates be formed in a liquid?

11. How would you endeavour to obtain large crystals of alum?

12. Describe an experiment to show the solubility of ether in water.

13. Sand and salt are stirred up in a bottle containing water in which some gas is dissolved. What methods would you use to recover the sand, salt, and gas from the water?

14. Give examples of the solvent power of water. How could you determine whether a sample of water contained dissolved solid matter?

15. Describe the apparatus you would employ to obtain water free from dissolved substances.

16. State in general terms the differences observable between a mixture and a compound of any two substances.

17. Explain the terms Mixture and Compound. How could you show that copper and sulphur (or iron and sulphur) can exist together, either as a mixture or as a chemical compound?

CHAPTER XV

THE ATMOSPHERE

Chemical Action.—In the study of Chemistry it is advisable to start with common and well-known substances, and to examine these as far as possible. The knowledge so obtained is then of service in the further study of other and less common substances. First, therefore, from its importance and its universal presence, we shall investigate the atmosphere. The physical properties have already been studied. We know it to be a gas possessing weight and capable of exerting pressure (p. 108). We have now to study its chemical properties. What is it composed of? How does it behave towards other substances? To answer these questions it will be necessary to carefully study the changes which different substances undergo when exposed to the influence of the atmosphere, selecting at first those changes which appear simplest. Of these the rusting of metals—say, iron—may be chosen. The first problem to be solved is, what is this rusting? Does the iron lose, or give up, anything? Or, on the contrary, does it gain anything? To answer these questions the following experiment should be done.

> Expt. 209.—Weigh carefully a watch glass with some iron filings or tacks, add a few drops of water (since we know that iron rusts best in the damp) and allow it to stand. At the end of a few days warm gently to drive off the water, and when *quite dry* again weigh, and note carefully the weight. Has the iron gained or lost?

If the weights have been carefully taken, it will be found that the iron has *gained*, and the experiment has furnished the necessary answer, so that we may write:—

Iron gains in weight during rusting.

Now we must ask—Did the gain come from the water we added, or from the atmosphere? To answer this question it will be necessary to allow the iron to rust in a closed space, so that if anything is taken from the air we can notice it.

EXPT. 210.—Place some iron filings in a muslin bag and tie the bag to a piece of glass rod. Moisten well (better dip in a solution of sal ammoniac) and place it in a bottle of air inverted over water (Fig. 110). Examine after a few days. It will be seen that the water has risen in the glass, showing that some part of the atmosphere has been abstracted by the iron in rusting.

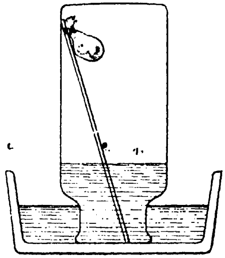


FIG. 110.—Action of Iron Filings on Air.

EXPT. 211.—Tightly place your hand on a card under the mouth of the jar so as to allow no water to escape, set upright and place a burning taper into the jar. Note what happens, but do not throw away the water.

The taper being extinguished proves to us that the gas left in the jar differs from air in the respect that it does not allow substances to burn in it. Hence the material taken away from the air by the iron, and which with the iron formed rust, is that part of the air which is concerned in burning, and we may state:—

Iron in rusting gains weight, taking some material from the air, and this material is the part of the air concerned in burning.

EXPT. 212.—Next measure in a graduated vessel the quantity of water in the bottle. This is equal to the quantity of gas which has been used by the iron. Also measure the quantity of water the bottle holds. This is the quantity of air it originally held.

Chemical Composition of Air.—It will be found by repeating the experiment with different bottles at different times, and with different quantities of air, that the ratio of these two volumes is always roughly 1 : 5, and is independent of the size of the bottle,

&c. Hence it is clear that this gives us the proportion of the *active part* of the air, and we may sum up the result of the experiment as follows.

Air contains 80 per cent. *inactive part*, which does not allow things to burn in it, so that the remaining 20 per cent. must be the part concerned in burning, and which we have called the *active part*. To the *inactive part* we shall henceforth give the name **Nitrogen.**

This gas nitrogen is an example of a very inert element, since it can only with difficulty be made to combine with any other element. It does not burn, nor will it allow other substances to burn in it; or, as usually expressed, it *does not support combustion*. It can be made to unite with a few metals, as magnesium or lithium. If either of these be heated with nitrogen the gas is absorbed and a compound of magnesium and nitrogen called magnesium nitride, or lithium and nitrogen called lithium nitride, is obtained.

If the nitrogen obtained in Experiment 210 be thus heated with one of these metals it is found that about 1 per cent. of it remains unabsorbed. This residue is another substance present in the atmosphere to the extent indicated, and called **Argon.** Argon is also very inert, more so even than nitrogen. Owing to this, even until the year 1894, its presence in the air had been completely overlooked, although, nearly a century before, the eminent chemist Cavendish had unknowingly obtained some, regarding it as an impurity which he had overlooked—an example of the importance of giving attention to the minutest details in scientific investigations.

The Active Constituent of Air.—We must now endeavour to obtain and examine the active part of the air which disappeared during the rusting of the iron. Knowing that it is now present in the rust, the most evident plan would be to endeavour to obtain it from this source. The ease with which iron rusts, that is, the readiness with which it takes up the active part of the air, should indicate that it would probably be very difficult to obtain it from this source, and that some other rust which is more difficult to prepare would probably be better for our purpose. The most convenient is the rust of mercury, which is a red powder not easily formed. It can be obtained if mercury is heated for a considerable time in air, when it forms as a red scum upon the surface of the metal. The changes which occur when this rust is heated should be carefully described in a notebook.

> EXPT. 213.—Place some of the mercury rust (known as red oxide of mercury) in a tube of hard glass closed at one end, and heat strongly. Notice the darkening of the powder, also the dark deposit which collects round the inside of the tube above the powder. Place into the tube a splinter of wood which has been just extinguished and is still glowing. Note that it glows more brightly, or even bursts into flame. Allow the tube to cool, and notice that the powder returns to its original colour. With a piece of wood or glass scrape off the dark deposit from the walls of the tube. It is seen to be bright metallic quicksilver or mercury.

What has this experiment taught? It has shown that by heating the rust of mercury we obtain mercury itself, and a gas in which wood burns more brightly than in air. Is this what we should expect to find supposing we have obtained the *active part* which we hoped to get? It certainly is, for we have found that the gas occupying the other four-fifths of the air does not support combustion at all, and thus acts as a diluent. Consequently the active part itself would be expected to support combustion very vigorously.

Absolute proof is, however, forthcoming. It may be proved, by weighing, that the weight of the original mercury is equal to that left after the experiment, provided that all the rust is decomposed and no mercury is lost. This shows that the gas escaping from the rust is the same gas as that taken from the air. This gas is called **Oxygen**, and the change may be thus stated:—**Oxide of mercury, when heated, decomposes into mercury and oxygen.**

Preparation and Properties of Oxygen.—As the quantity of oxygen obtained by the above method is comparatively small, and the oxide is expensive, a more convenient source of the gas is a white crystalline powder called Potassium Chlorate.

> EXPT. 214.—Place a little potassium chlorate in a test-tube and heat. Observe that the mass crackles, melts, and gives off a gas. Test by a glowing match, and see that the gas is *oxygen*. The gas is, however, given off more readily, and without fusion, if a little manganese dioxide is added to the chlorate. By this means we can collect some jars of the gas.

EXPT. 215.—Fit a delivery tube *B* to a flask *A*, and in the flask put a mixture of potassium chlorate and manganese dioxide. Let the end of the delivery tube dip under water contained in the trough *C*. Fill five jars with water and invert them in the trough. Heat the flask, and when the gas is being evolved place the mouth of the jar *D* over the end of the tube so that the gas displaces the water. When all the jars are full take out the delivery tube from the water, and leave the flask for further examination. We must now examine the gas. Its action upon burning wood has been seen before.

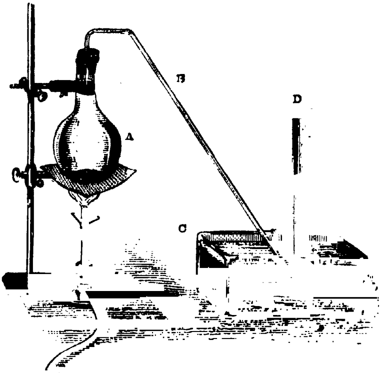


FIG. 111.—Preparation of Oxygen.

EXPT. 216.—*First jar.*—In this plunge a piece of red-hot charcoal placed in a deflagrating spoon (a small upturned iron spoon with a long handle) (Fig. 112). Note the brilliancy of the combustion. Now pour into the jar some clear lime-

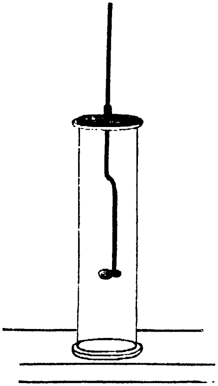


FIG. 112.—Jar with Deflagrating Spoon in it.

water (*i.e.* some of the clear liquid which is formed if lime and water be shaken together and allowed to stand. It is really a solution of lime in water). Notice that it turns milky owing to a white powder being diffused through the liquid.

EXPT. 217.—In the *second jar* burn some phosphorus contained in the deflagrating spoon. Note the great brilliancy of the combustion and the dense white fumes. Add water and shake—the fumes dissolve. In the solution put a little blue litmus paper. Observe that it is turned *red*.

EXPT. 218.—*Third jar.*—Perform a similar experiment with sulphur. There are few fumes, but a strongly smelling gas is obtained, also soluble, turning the blue litmus red.

EXPT. 219.—In a dry deflagrating spoon put a small piece of the metal sodium,[1] light it and put it in the *fourth jar* of the gas. Observe the fumes formed. Dissolve these in water and try the effect of the solution on litmus paper. It does not turn the blue colour to red but has exactly the reverse effect, and changes the red litmus blue. Feel the water, it has a peculiar soapy feel.

EXPT. 220.—Obtain a piece of iron wire (a thin steel watch spring will do), and dip one end into a little melted sulphur, and when the sulphur is burning place the wire in the *fifth jar* of oxygen. Observe that the sulphur burns and also starts the combustion of the iron which continues to burn with a brilliant shower of sparks. After the burning has ended, observe that a quantity of an insoluble solid (iron rust) has been formed.

Oxides and their Properties.—In all these cases of burning, both the substance burnt and the oxygen disappear and a new product is found. They have in fact combined to form a compound which we term an *oxide*, the name given to the compound of oxygen with some other element. We have noticed also that the oxides differ completely in their action upon litmus, for whereas some turn the blue litmus red, others do exactly the reverse. Substances which, like the solution of the oxide of phosphorus or sulphur, change the blue litmus red are **acids**.

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[1] Great care must be taken when using sodium, which must never be allowed to touch damp materials. It is kept under naphtha until used, and should never be handled with the fingers. It should be dried by blotting paper when taken from the bottle and cut with a clean knife, the pieces not used being immediately replaced in the bottle.

Expt. 221.—The student may verify this by trying the different *acids* in the laboratory.

Acids are further characterised by a sour taste, even when in dilute solution, and by other far more important chemical properties. The solution of oxide of sodium is what is called an *alkaline* solution. The student will do well to note that, as a general rule, he may regard the oxides of metals as different in their properties from the oxides of the non-metallic elements.

We may now write :—

Iron and oxygen form oxide of iron.
Phosphorus and oxygen form phosphorus pentoxide.
Carbon and oxygen form carbon dioxide.
Sulphur and oxygen form sulphur dioxide.
Sodium and oxygen form sodium oxide.

The meaning of the term *di*, &c., will be understood later.

We must now return to the *flask which contains the residue of the heated mass.*

Expt. 222.—Heat well, till no more oxygen is given off. Boil up the mass with water and filter it. The manganese dioxide remains unchanged on the filter paper. Evaporate the solution to dryness, and there remains a white solid which is not potassium chlorate and which tastes something like salt.

This substance is called potassium chloride, and we may therefore write :—**Potassium Chlorate, when heated, decomposes into potassium chloride and oxygen.** The manganese dioxide did not change ; what, then, was its rôle ? This question is one which was for a long time explained by calling the manganese dioxide a *catalytic agent*, and such actions, viz., those accelerated by the presence of a substance, which itself does not change, were called *catalytic actions.* In reality, however, the manganese dioxide goes through a series of changes, but the final one leaves it in its original condition.

Quantitative Character of Chemical Action.—The heating of potassium chlorate and manganese dioxide must also be done quantitatively, as the numbers obtained will be of use to us in our later work.

Expt. 223.—Procure a test-tube and put in it a little manganese dioxide, and place near the mouth a plug of glass wool; weigh. Add potassium chlorate and put back glass wool; weigh again. Connect up as in Fig. 113 and heat until no more oxygen is evolved; measure the volume of gas given

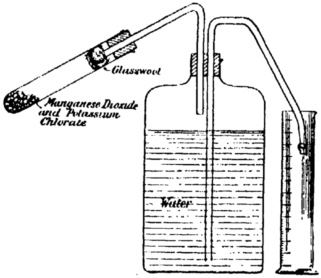

Fig. 113.—To Illustrate Experiment 223.

off, that is the volume of water driven out of A. Allow the tube to cool and weigh again.

Enter your results thus:—

	Wt. of tube, manganese dioxide, and potassium chlorate	36·85 gms.
	Wt. of tube and manganese dioxide	35·32 gms.
1.	Wt. of potassium chlorate	1·53 gms.
	Wt. of tube, manganese dioxide, and potassium chlorate	36·85 gms.
	Wt. of tube, manganese dioxide and potassium chlorate, after heating	36·25 gms.
2.	Wt. of oxygen evolved	·60 gms.
3.	Volume of oxygen	430 c.c.

We thus find, if the experiment is carefully performed, that from 100 gms. of potassium chlorate we should obtain $\frac{100}{1\cdot53} \times \frac{\cdot60}{1}$ gms., *i.e.*, 39 gms. of oxygen, and, further, that 430 c.c. of oxygen weigh ·60 gms. that is, 1 litre weighs 1·39 gms. We have not here, however, corrected the volume of gas for the pressure and temperature, and this should be done, the barometric pressure and air temperature being read at the time of the experiment, and the volume of gas corrected for the pressure 760 mm. and temperature 0° (page 112). The weight of 1 litre of oxygen will be then found as slightly over 1·4 gms.

Summary of Results.—Air contains about 80 (more accurately 78) per cent of an inert gas, nitrogen, in which things cannot burn. About 20 (more accurately 21) per cent. is a very active gas, *oxygen*, which supports combustion very vigorously. Substances burn in it with the formation of compounds called *oxides*. Some of these oxides dissolve in water and form *acids*, while others have properties essentially different from the acid-forming oxides. A little less than one per cent. of the air consists of a very inert gas named Argon.

Various substances present in Air.—Besides these, which are the chief substances in air, there are others which are present to a small extent. The presence of some of these may be indicated by the following experiments:—

EXPT. 224.—Into a *large* glass flask pour a few c.c. of clear lime-water. Close the jar and shake vigorously. Examine the lime-water.

EXPT. 225.—Leave a few pieces of dry calcium chloride exposed in a saucer to the air. Observe after a few hours.

EXPT. 226.—Leave a weighed quantity of sulphuric acid exposed to the air in a beaker for a few days. Again weigh.

EXPT. 227.—Put some ice (if not available, a little sodium sulphate and hydrochloric acid) into a small beaker and watch the outside of the beaker.

In the first of these experiments we find that the lime-water becomes tinted and milky. This does not happen with oxygen alone (try it) or nitrogen, but we have already seen that the milkiness is produced by the gas produced by burning carbon, *i.e.* carbon dioxide. This gas is, in fact, present in the atmosphere, but only to a small extent, viz., about ·04 per cent., that is, 4 parts in 10,000. Even this minute quantity is, however, very important, as plants derive the carbon which enters into their composition from this source. Their green parts, under the action of sunlight, decompose the oxide, and give off the oxygen.

If, instead of atmospheric air, air expired in breathing be used in the above test, it will be seen that the quantity of carbon di-oxide is much greater, indicating that this is produced as a product of the animal economy, resulting, in fact, from the union of the atmospheric oxygen with the carbon present in animal tissues. Owing to this all animals require oxygen, and soon expire if deprived of it. Fish and many other marine animals are thus entirely dependent, for their supply of oxygen, upon the air which is dissolved in the water.

The experiments (225-226) indicate the absorption of moisture by the calcium chloride and sulphuric acid, and it is this moisture which is present in the air that becomes deposited upon cold objects in the form of dew, as in Expt. 227.

The quantity of moisture which is present in the air in the form of water vapour varies considerably from time to time, being as a rule greater in hot weather than in cold.

In the neighbourhood of towns, also, there are always present small quantities of compounds which are formed in works, &c. These include compounds of ammonium, oxides of sulphur, &c. Large quantities of dust particles and minute bacteria are also found in air, the number of these little organisms varying from one or less to many thousands per cubic centimetre.

Air Compared with Oxygen.—We have now seen that when a substance burns in oxygen a new compound is formed, which is an oxide. But we must go further, and ask—Is this also the case when it burns in air? Does it take away the oxygen and form an *oxide*, leaving the nitrogen unacted upon, as we should expect? This can be readily tested.

Expt. 228.—Put a small piece of dry phosphorus in a test-tube, cork the tube tightly, and warm. Note that the phosphorus ignites and forms white fumes, as when it burns in oxygen. But the flame is not so bright. When cool again, take out the cork under water. Note what happens. Put back the cork and shake up, test the remaining gas, and test the water with litmus. Measure the volume of the water, *i.e.* the gas abstracted, also the volume the tube holds.

As was the case in rusting, one-fifth of the air has been used, this being what we found was the quantity of oxygen present; the white fumes formed, too, combine with water to make an acid, as when the phosphorus burned in oxygen. In fact, all tests prove the products to be identical. Hence, in air, also, the burning consists of the union of the burning substance with the oxygen of the air to form a new compound—*an oxide.*

It will be easy for the student to verify this by burning in air the different substances which were burnt in oxygen, and carefully comparing the products.

Gain of Weight during Burning.—If the explanation given above be correct, it is evident that substances should gain in weight when burnt, *i.e.* the products of combustion should weigh more than the original substances. Do they? We may test this, for, say, phosphorus, sulphur, and magnesium wire.

EXPT. 229.—Weigh a crucible and lid and a piece of magnesium, which, folded lightly, is placed in the crucible. Heat strongly in a burner, taking care to let no fumes escape. To do this, keep on the lid, and only raise a little when the flame is removed. The magnesium is seen to burn brightly in places; but, if care is taken, no fumes are lost. When finished, the whole mass should be in the form of a white powder. Allow to cool, and weigh the crucible with the lid and powder. Subtract the weight of the crucible and lid to find the weight of the powder. Is it more or less than the weight of magnesium? It will almost certainly be found to be *more*, and, if the experiment is carefully done, it will be found that the magnesium has gained about 66·5 per cent.

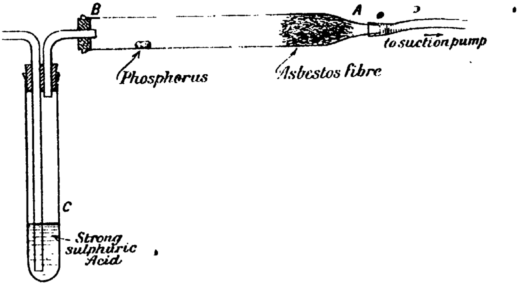

FIG. 114.—To illustrate Experiment 230.

EXPT. 230.—Get a hard glass tube, *BA* in figure, pack the drawn-out end with asbestos fibre; weigh carefully. Place in it a small piece of dry phosphorus (about ·2 gram) and again weigh. Connect up as shown in figure, the end *A* being attached to a suction-pump which can be fitted to a tap. This draws a current of air through the tube, and, in order to dry the air, it is first made to pass through the test tube C, containing strong sulphuric acid. Warm the phosphorus slightly —it ignites. Take away the burner and allow it to burn, the fumes being stopped by the asbestos. A quantity of red deposit is also found (another form of phosphorus), and this may be got rid of by strongly heating. When cool, disconnect, and again weigh the tube. It is found, if carefully done, that the phosphorus gains about 128 per cent.

Expt. 231.—If we desire to do the same with sulphur, we must alter our apparatus, as the product is a gas. This is readily done. The tube BA, as before, contains the sulphur, but the air together with the products of combustion are drawn through a strong solution of caustic potash in the test-tube D, the air having been previously drawn through the tube C containing strong sulphuric acid (Fig. 115). The sulphur dioxide is absorbed in D, so that if this tube be weighed before and after the experiment we obtain the weight of the gas, while the weight of the tube AB with sulphur in it (1) before heating (2) after heating gives the weight of sulphur which has burnt away. From these results calculate the increase of weight for 100 grms. of sulphur. It is seen that the sulphur gains 100 per cent.

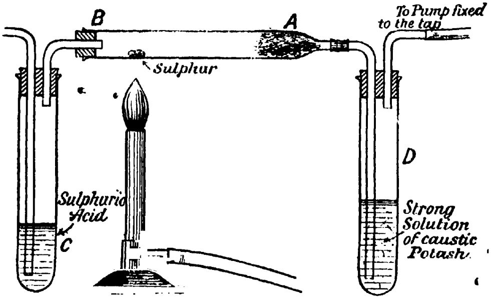

Fig. 115.—Burning Sulphur in a Current of Air.

Combustion and Rusting.—A comparison of these experiments with those on the rusting of iron should be made in order to plainly see that combustion and rusting are essentially similar actions, inasmuch as both are the formation of *oxides*, the only difference being that in burning this is attended with the production of *light*. How is this light brought about? It is owing to the heat consequent on the chemical union. When two substances unite to form a new compound, heat is generally evolved, and this may be sufficient to raise the substances and the products of the combustion to such a temperature that they become luminous. If gases are present among the products of combustion they may become luminous and pass away as *flames*, which hence consist of *gases, produced by the combustion, raised to such a high temperature that they become self-luminous.*

If a jet of combustible gas be allowed to burn in air, it is clear that combustion can only occur where the gas and air are in contact, *i.e.*, on the outside of a jet, hence flames are generally *hollow*, as may be seen by placing a piece of wire gauze over the flame of a candle or Bunsen burner (Fig. 116).

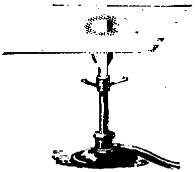
FIG. 116.—Hollow structure of a Flame.

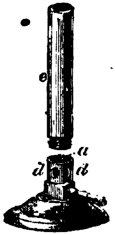

FIG. 117.—Bunsen Burner.

EXPT. 232.—Place one end of a piece of glass tubing into the *middle* of a Bunsen burner flame, or a candle flame. Slope the tubing upwards away from the flame, and apply a light.

It will be found that the interior of the flame consists of unburnt gas capable of combustion, which, from what has been previously said, is what we should expect.

EXPT. 233.—Light a Bunsen burner and examine the flame. Hold in it for a short time a piece of glass tubing. Now close the air holes, *d*, at the base of the burner (Fig. 117) and observe the change in the appearance of the flame. Hold in the flame a piece of glass tubing, and observe what happens.

It is thus found that closing the air supply at the base of the burner greatly increases the luminosity of the flame, while also soot (or carbon) is deposited on a glass rod held in the luminous flame. The increased luminosity appears, therefore, to be due to the presence of solid carbon, which becomes white hot or *incandescent*, owing to the high temperature of the flame. In general, also, when solids are present among the products of combustion the flame becomes intensely luminous, owing to the **incandescence** of the solid, whereas if the products are entirely gaseous the flame is as a rule not very luminous. Solids raised to a white heat are largely used for illuminating purposes. The *incandescent* electric lamps consist of a filament of carbon in a vacuous globe raised to a white heat and made incandescent by means of the electric current. It is important to notice that here there is *no* burning or combustion, since no oxygen is present. The *limelight* is produced by the incandescence of a block of hard lime, which is made white hot by the flame obtained by burning a mixture of coal gas and oxygen. In the *incandescent* gas light a network is formed of some white incombustible oxide (the *mantle*), and this is heated by the flame of what is practically a Bunsen burner, and so made white hot. It is similarly important to notice that in the last two cases the lime and mantle do not burn, but are really heated by the burning gases.

It must be emphasised, however, that although the presence of incandescent solid particles is a frequent and important cause of luminosity of flames, yet it is not the sole cause, as some flames are intensely luminous, although no solids are present. Many factors, indeed, influence the luminosity, as, for example, the temperature and the density of the gases.

Oxidising and Reducing Flame.—It has been already seen that in the interior of the flame hot combustible gases are present which are capable of burning, *i.e.*, of combining with oxygen, and which can thus abstract oxygen from oxides and other oxygen compounds. It is said to be the **reducing** part of the flame. At the top apex of the flame, however, no combustible gases are present, as all have been burnt, but there is present a quantity of strongly heated air which can readily *give up oxygen*. This part may be hence said to be the **oxidising** part of the flame.

CHIEF POINTS OF CHAPTER XV

Iron gains in weight during rusting, taking some material from the air, and this material is the part of the air concerned in burning. These facts were found out and established by the series of experiments which the reader has, it is presumed, either performed or seen performed.

Nitrogen is the inactive part of the air. It makes up 80 per cent. (roughly) of the atmosphere. It is a very inert element. It does not burn, nor will it allow other things to burn in it. It can be made to unite with a few metals, such as magnesium and lithium, forming *nitrides*.

Argon is another very inert element, which was discovered in 1894. It makes up about 1 per cent. of the air.

Oxygen is the active part of the air. When red oxide of mercury is heated it decomposes into mercury and oxygen. This gas can also be obtained by heating potassium-chlorate. The oxygen is given off from this substance more readily if it is mixed with manganese dioxide before being heated. The substance left after driving the oxygen out of potassium chlorate by heating it is called potassium chloride and is very like common salt.

Oxides.—When some elements are heated in oxygen they unite with it, forming **oxides**. Thus—

Iron and oxygen form Oxide of iron
Phosphorus and oxygen form Phosphorus pent-oxide
Carbon and oxygen form Carbon di-oxide
Sodium and oxygen form Sodium oxide.

Some oxides unite with water to form **acids** which turn blue litmus red ; the oxide of phosphorus is an example.

Other oxides unite with water to form an **alkaline solution** which has a soapy feel ; the oxide of sodium is an example.

Quantitative Character of Chemical Action.—It has been found that when the experiment of heating potassium chlorate is carefully performed a given quantity of this substance always gives the same quantity of oxygen.

Various Substances present in the Air.—*Carbon dioxide* is generally present to the extent of 0·04 per cent., or 4 parts of carbon dioxide to 10,000 of air. This carbon dioxide is a very important part of the food of plants. Its presence is partly due to the breathing of animals.

Moisture is always present. Its amount varies with the temperature. It is often deposited as dew.

Ammonium compounds, oxides of sulphur, dust particles, and *bacteria* are to be found in the air in the neighbourhood of towns.

Gain of Weight during Burning.—It has been seen by performing experiments with phosphorus, sulphur and magnesium, that there is always a gain in weight when a substance is burnt. Thus, phosphorus gains about 128 per cent., sulphur 100 per cent., and magnesium 66·5 per cent.

Combustion and Rusting.—These are essentially similar. The only difference is that combustion is attended with the production of **light**, due to the high temperature to which the products are raised by the heat of the chemical action.

QUESTIONS ON CHAPTER XV.

1. Describe experiments which prove that air is composed of at least two gases.

2. How would you show by experiments that only one part of the air is concerned in (*a*) burning, (*b*) rusting?

3. Describe experiments indicating that rusting and burning are essentially similar operations.

4. Describe briefly the sequence of experiments which indicate the existence of oxygen in air, and give a mode of preparation of this gas.

5. How is oxygen most conveniently prepared? Give an account of its characteristic properties.

6. What are oxides, and how may they be obtained?

7. How can nitrogen be obtained from the air, and what are its chief properties?

8. By what experiments may the presence of (*a*) aqueous vapour, (*b*) carbon dioxide, in the air be indicated.

9. Describe experiments which prove that during rusting and burning an increase of weight occurs.

10. How would you find the quantity of oxygen and nitrogen present in a given volume of air?

11. By the combustion of different substances various oxides may be produced. Point out how by the action of water these oxides may be divided into two classes, and give examples.

12. Describe (*a*) a direct, (*b*) an indirect method of finding the weight of a litre of oxygen.

13. What is argon? How has it been obtained?

14. To what are flames due; what is a reducing flame, and what is an oxidising flame?

15. On what does the *luminosity* of a flame depend?

16. The gas obtained by burning a mixture of coal gas and oxygen is not very luminous. How may it be utilised as a source of intense light?

17. What is meant by incandescence? What is the cause of the incandescence in (*a*) an electric lamp, (*b*) an incandescent gas light?

18. What is the difference between an ordinary flame and that of a Bunsen burner?

19. Oxygen and nitrogen are the two chief ingredients in common air. State reasons for the usual belief that they are not combined together chemically.

CHAPTER XVI

WATER AND THE ELEMENTS COMPOSING IT

Water.—We must next turn our attention to *water* and endeavour, as we did in the case of air, to investigate its chemical composition, &c. Let us briefly review what we already know of it.

It is a clear liquid with a blue-green colour, which is best seen by causing light to pass through a considerable length of water. It boils at 100° C., and is then converted into steam. It freezes at 0° C., becoming ice. Its density, *i.e.*, mass of 1 c.c., is 1 at 4° C. and slightly less at any other temperature, owing to the fact that water expands when either cooled below or heated above 4° C. It has the power of dissolving many substances, *e.g.*, salt, sugar, &c., forming solutions from which the water may be evaporated away, leaving the solid behind. But these facts tell us nothing of the chemical nature of water because we have no changes in composition. The water, present all the time, is not converted into any new product. We must therefore try to get a new product from the water. We will first try the action of metals, and as in the case of our studies upon air, we will use iron first. If iron be left in water it is found to form a considerable quantity of *rust*; but this may be due to the air which we now know may be dissolved in water; and if the iron be placed in water in a tube, this water, having been first well boiled to drive off the air, and the tube then sealed in a blow-pipe, the iron either does not rust, or does so to only a very slight extent. There is, therefore, no conclusion derivable from this experiment. Perhaps, however, if the iron were heated it might act on the water.

Expt. 234.—Place some iron filings in the tube CA, and let the end A dip under water. To the end C fit a delivery tube from a flask containing water previously boiled to drive off the dissolved air. Heat the iron filings well, and boil the water in the flask so that the steam passes over the heated iron and then into the water, where it condenses. Now place over the end A an inverted test-tube of water, and note that the steam is not completely condensed, but that minute bubbles ascend to the top of the test-tube. When you have so obtained a sufficient quantity of gas (half a small test-tube), first remove C from the flask of water; then stop the boiling. Close the end of the test-tube with your thumb, and holding a lighted match to the mouth, open the tube. Observe that the gas *burns*. Examine the iron filings in the tube CA, and see that a quantity of rust has been formed.

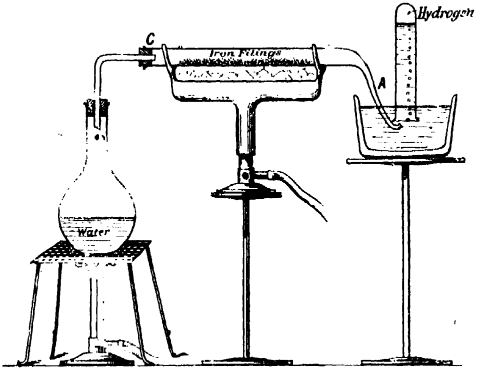

FIG. 118.—Action of Steam on Heated Iron.

Action of Sodium upon Water.—We see, therefore, that from the steam and iron we obtain iron rust, that is, iron oxide, and an inflammable gas. We must investigate this further, as it appears to show that the water contains this inflammable gas and oxygen. We must find something which has a more powerful action on oxygen than the iron. Such a substance is the metal sodium, which we have before used.

EXPT. 235.—Place a small piece of sodium in water in an evaporating basin, and observe the action. Feel the water left after the sodium has all disappeared, and test it with litmus. Evaporate away the water. Note the residue.

In this case we could only see that the action was energetic and a new product was formed, while the soapy feel of the water, and the action on litmus, appeared to indicate that this product was that previously obtained when sodium burned in oxygen and the fumes were dissolved in water. We could not, however, see whether any gas came off, as the experiment was done in the open air and the sodium floated on the surface of the water. We must therefore try to alter our experiment so as to collect any gas which may be evolved.

Expt. 236.—Place the sodium in a small piece of lead tubing, the ends of which are nearly closed, and drop into water (or place down over the sodium a thimble with holes in the top). A gas is seen to come off. Collect this in an inverted tube full of water, and by this means obtain three test-tubes of the gas, and observe that it is colourless and odourless.

Expt. 237.—Take out two tubes and hold them for the same time, say, 30 seconds, (1) with mouth up, (2) with mouth down. Then apply a light to the mouth of both.

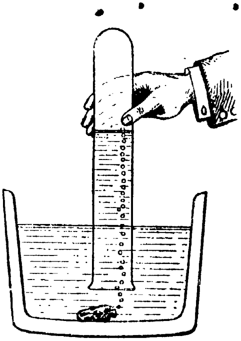

Fig. 119.—Action of Sodium on Water.

The slight explosion of (2) shows it to contain an explosive gas, while the absence of any effect with (1) shows that the gas has disappeared. Hence it is seen that the gas escapes from a tube held mouth upwards, but not from one held mouth downwards. It is therefore lighter than air, being in fact the lightest gas known.

Expt. 238.—Now try the 3rd tube, holding it mouth down, and place a lighted match up into the tube. Note that the gas does *not* explode, but burns quietly, while the match is extinguished (Fig. 121).

Preparation and Properties of Hydrogen.—From water and sodium we have therefore obtained an inflammable gas, lighter than air, which does not support combustion, while there is also formed a solution which behaves like the solution of sodium oxide. The most natural inference is that the water contains this inflammable gas which we call **hydrogen**, and oxygen. Before proceeding to verify this, which we must do, it will be well to examine more carefully the properties of the inflammable gas, and to do this we should collect it in greater quantity than hitherto. For this purpose we must act upon a metal with a dilute acid instead of water.

EXPT. 239. Select a flask and fit it up as is shown in Fig. 120. Be very careful that the stopper and the tubes respectively fit very closely. Into the flask put enough granulated zinc to cover the bottom. Pour some water upon the zinc. Arrange the delivery tube in the trough as you did when you were making oxygen. Pour a little sulphuric acid down the thistle-headed acid funnel, and be quite sure that the end of the funnel dips beneath the liquid in the flask. Do not collect bottles of the gas until you are sure pure hydrogen is being given off, which you can find out in this way. Fill a test-tube with water and invert over the end of the delivery tube. When it is full of gas, still holding it upside down, take it to a flame (which should not be near the flask you are using) notice that there is a slight explosion. Continue this until the hydrogen burns quietly down the test-tube. When this happens you may proceed to fill one

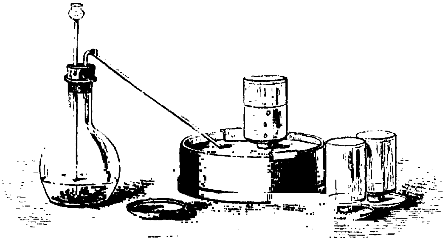

FIG. 120.—Preparation of Bottles of Hydrogen.

or two bottles. When the bottles have been filled, it is better not to remove them from the water until you want to use them. Collect also a soda-water bottle half full of the gas.

EXPT. 240.—Test one jar of the gas by means of a lighted match or taper as in Expt. 238. Observe that the gas burns around the mouth of the jar but that the taper is extinguished, but on being taken out, again becomes alight on passing through the flame of the burning hydrogen (Fig. 121).

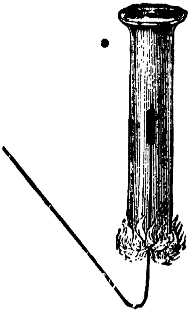

FIG. 121.—Hydrogen burns, but extinguishes a Flame.

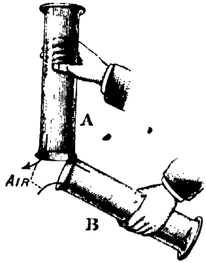

FIG. 122.—Illustration of Low Density of Hydrogen.

EXPT. 241.—Take a full jar of the gas and hold it mouth upwards below a second smaller jar held mouth downwards as shown in Fig. 122. On testing with a lighted taper observe that the gas has left the lower jar and filled the upper. Many experiments, as the filling of balloons or soap bubbles, may also be performed to demonstrate the extremely low density of hydrogen.

EXPT. 242.—Wrap your hand well in a duster and with it hold the soda-water bottle. Take it out of the water so that the water runs out, and the bottle is now filled with a mixture of hydrogen and air. Apply a light and you will not fail to observe that an explosion results.

EXPT. 243.—Apply a light to the end of the delivery tube and allow the hydrogen to burn. Observe that it burns with a pale blue flame, which after a time becomes yellow. As this colouration does not occur however until the glass becomes hot, we must regard it as due to the glass.

We have now a means of obtaining hydrogen in considerable quantity, and have observed that it is a colourless, odourless gas, considerably lighter than air, which does not support combustion, but which burns itself, and which with air forms a highly explosive mixture.

Production of Oxide of Hydrogen.—We must next endeavour to obtain and to examine the compound which is produced by the burning of hydrogen, that is, the oxide of hydrogen.

EXPT. 244.—Arrange a flask as before for the production of hydrogen. Pass the gas through a tube containing chloride of calcium in order to thoroughly dry it. Allow it to burn under a retort which is kept cool by a stream of water flowing in at the tubule and out at the end of the neck (Fig. 123). Observe the formation, on the side of the retort, of a clear liquid which collects and drops into a beaker placed to receive it. By this means sufficient of the liquid can be obtained to identify it, especially if a number of students all work together and add the liquids.

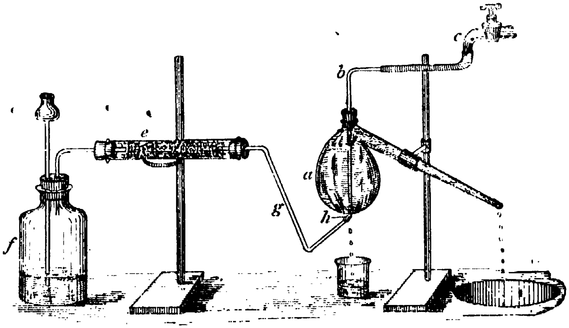

FIG. 123.—Production and Collection of Oxide of Hydrogen (Water).

EXPT. 245.—Take the density, freezing point (a mixture of sodium sulphate and hydrochloric acid forms a very convenient freezing mixture), and boiling point of the liquid. You will find these are 1, 0° C., and 100° C. respectively, a result which is sufficient to enable us to state that the liquid is identical with pure water.

We have hence completely proved our previous ideas to be correct, for we had been led to believe that water contained hydrogen and oxygen. We have found it is an oxide of hydrogen. Hence we may now say :—**Hydrogen in burning produces water which is, therefore, an oxide of hydrogen.**

We should next examine the flask itself which we used for the preparation of the hydrogen in order to see whether any other products have been formed, and to find out what has become of the zinc.

> EXPT. 246. Filter off the liquid in the flask from the undissolved zinc (sufficient zinc should be used to leave a quantity still undissolved ; if all has disappeared add more and wait till the action ceases). Partially evaporate the liquid and allow it to crystallise. You will find that a quantity of clear colourless crystals are formed. Examine them and sketch the most perfect. Heat some of the crystals in a tube and observe that they melt, give off water (which can be collected and proved to be water), and leave a white powder.

These crystals are a compound formed from the zinc and part of the sulphuric acid, and are known as *zinc sulphate*. Many crystals behave similarly on heating, *i.e.*, lose water which they previously contained, and become converted to a powder. The water contained in a crystal and evolved on heating is known as *water of crystallisation*. We may therefore state **sulphuric acid and zinc form hydrogen and zinc sulphate.**

Proportions of Oxygen and Hydrogen in Water.—We must next endeavour to find the proportions in which the oxygen and hydrogen combine during the formation of water. We may do this in either of two ways, viz., by finding the weight of the gases or by finding the volume of the gases which combine. For the latter it is necessary to measure out definite volumes of oxygen and hydrogen, cause them to combine, then measure the volume of gas which remains uncombined and ascertain which gas it is. This is usually done in a piece of apparatus known as an **Eudiometer** (Fig 124.). In its simplest form this consists of a long glass tube closed at one end and graduated in

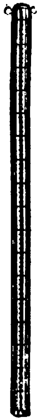
FIG. 124.—Eudiometer.

equal volumes, usually cubic centimetres, by divisions marked on the glass. Through opposite sides of the tube at the closed end pieces of platinum wire are passed and fused into the glass, being so arranged that they do not quite touch one another. Outside the tube the platinum wires are bent into loops to which wires from an electric coil may be attached.

Composition of Water by Volume.—To use the eudiometer it is first completely filled with mercury and inverted over more mercury contained in a trough. A suitable quantity of pure dry oxygen is then bubbled into the tube and the volume (after the necessary corrections for temperature and pressure) is recorded. Pure dry hydrogen is next bubbled into the tube, using a considerable excess of one or other of the gases. The volume is again recorded (with necessary corrections as before) and then, keeping the eudiometer firmly pressed upon a sheet of india-rubber, or felt, at the bottom of the trough, the gases are made to combine by causing an electric spark to pass between the platinum wires inside the tube. As soon as the spark passes, the two gases combine, with a flash of light. The eudiometer is slightly raised from the india-rubber (but of course not above the mercury in the trough), and it is seen that the volume of the gas in the eudiometer is less than before the explosion, and that there is a film of moisture upon the interior of the tube. The corrected volume is again recorded and the nature of the gas ascertained. We then find the volumes of the gases which have combined in the manner indicated below :—

Corrected volume of oxygen ... 12 c.c.
Corrected volume of mixed gases ... 50 c.c.
Therefore corrected volume of hydrogen = 38 c.c.
Corrected volume after explosion = 14 c.c.
Gas left ascertained to be Hydrogen.

Hence the 12 c.c. of oxygen united with (38 - 14), *i.e.*, with 24 c.c., of hydrogen, and we find this result always obtains, namely, **2 volumes of hydrogen combine with 1 volume of oxygen to form water.**

If instead of doing the experiment in the above manner the eudiometer tube be kept heated above 100° C. during the experiment (by surrounding it with a larger tube through which the vapour of some boiling liquid was passed), it would be further found that the steam produced from these 3 volumes of oxygen and hydrogen would only occupy 2 volumes.

Such a process as this, the formation of a compound from elements, or from simpler materials, is known as a **synthesis.** We might also find the required ratio by the **analysis** of water, that is by breaking it up into its components, which we can do by passing an electric current through it.

Expt. 247.—This may be done by means of an electric battery for generating the electric current and a **Voltameter.** The latter is most simply made by closing the bottom of a funnel by means of a tightly fitting cork through which pass two platinum wires with small plates of platinum attached to the ends remaining in the funnel (Fig. 125). Over these plates are supported two glass test-tubes, and the tubes and part of the funnel are filled with water to which has been added a little sulphuric acid, as otherwise it offers great resistance to the electric current. The wires from a battery of three or four cells are connected with the ends of the platinum wires, and as soon as the connection with the battery is complete, provided there is clean metal at every junction, bubbles of gas are seen to rise from each platinum plate and to ascend into the tube and displace the contained liquid. After the experiment has gone on for half-an-hour, the gases may

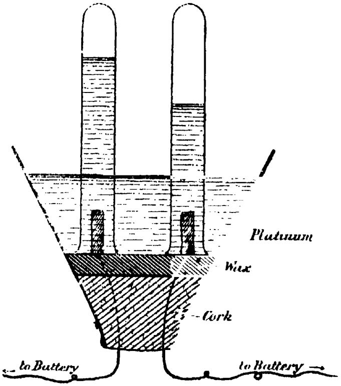

Fig. 125.—A Voltameter.

be tested and their volumes measured. It will be found that the volume of one gas is double that of the other, and that the gas of which there is the larger amount is *hydrogen* while the other is *oxygen*.

This experiment again proves what was found by synthesis, viz., that 2 vols. of hydrogen combine with 1 volume of oxygen, and further, it may be noticed that these experiments also prove that water is solely formed from these two gases, and contains no other constituent.

Composition of Water by Weight.—We must now endeavour to find the composition of water by weight, that is, the weights of oxygen and hydrogen which combine to form water.

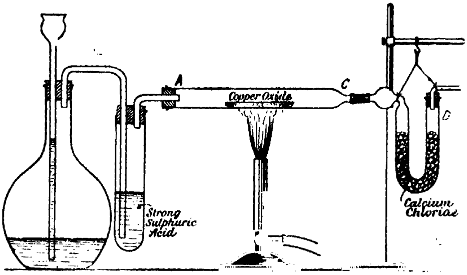

FIG. 126.—Composition of Water by Weight.

To do this it should be noticed that we only require the weigh of two out of the three substances concerned, *i.e.*, if we kn the weights of hydrogen and water (or of oxygen and water), t weight of the oxygen (or hydrogen) is readily calculated. T experiment is done by finding the weights of the oxygen a water, and for this it is best to use not oxygen itself, but so oxide which readily gives up its oxygen to the hydrogen, so t by weighing the oxide before and after the experiment we c ascertain the weight of oxygen which it has lost. The ox used for this purpose is oxide of copper, a black powder.

Expt. 248.—In the glass tube AC (Fig. 126) place a sm porcelain boat containing copper oxide, the boat and oxide be previously weighed. To the end C attach a U-tube containing calcium chloride, a substance which, we have already seen, readily absorbs moisture. At A attach a tube, as shown, by means of which dry hydrogen, as in Expt. 244, may be passed through the tube AC. The U-tube of calcium chloride must be carefully weighed before the experiment. When the hydrogen is passing, heat the tube AC, and observe that the black colour of the copper oxide changes to the red colour of metallic copper, and that water collects in the U-tube D. Take care that no condensed water remains in AC; if any is present, heat it and drive it over into D. Allow the tube to cool, disconnect, and weigh the boat of oxide and the U-tube again. Enter your results thus :—

Wt. of U-tube after experiment	36·473 grams
Wt. of U-tube before experiment	35·821 grams
Wt. of water produced	0·652 grams
Wt. of boat of copper oxide before experiment ...	5·562 grams
Wt. of boat of copper oxide after experiment ...	4·982 grams
Wt. of oxygen	0·580 grams

The weight of oxygen is 0·580 gram, and that of the water produced from it is 0·652 gram; therefore the weight of hydrogen must be 0·652 – 0·580 gram, that is, 0·072 gram. Write, therefore :—

Wt. of water	0·682 gram
Wt. of oxygen	0·580 gram
Wt. of hydrogen ...	0·072 gram

Therefore ·072 gram hydrogen combines with ·580 gram of oxygen, and hence 1 gram hydrogen combines with $\frac{\cdot 580}{\cdot 072} = 8$ grams of oxygen.

If the experiment is carefully performed this is the result always obtained, so that we have found that **Water is formed of eight-ninths its weight of oxygen with one-ninth its weight of hydrogen.**

This experiment, also taken in conjunction with the previous experiments upon the volumes of the gases, proves further that one volume of oxygen is 16 times as heavy as the same volume of hydrogen, so that 1 litre of hydrogen weighs $\frac{1\cdot4}{16}$, that is, nearly ·09 gram (accurately ·0896).

Natural Waters.—Water dissolves a larger number of things than any other liquid, and is in consequence of the greatest use to chemists. This explains why we cannot find pure water in nature. No sooner has the rain formed than it begins to dissolve various substances; in its passage through the air it takes up varying amounts of the constituents of the atmosphere, such as carbon dioxide and oxygen, and when the surface of the earth is reached the water dissolves out of the soil and the underlying rocks portions of all the soluble ingredients. The most soluble bodies are naturally dissolved to the largest extent. It will be seen later that the solvent property of water is considerably increased by the presence of the carbon dioxide it obtains in part from the air. When the amount of material dissolved in water is very great it gives a distinctive character to the liquid, which becomes known as a **mineral water.** Those natural waters which contain a compound of sulphur and hydrogen, called sulphuretted hydrogen, are spoken of as *sulphur-water*, if some compound of iron is the substance which has been taken up in large quantities, we have *chalybeate* waters formed. *Effervescent* waters have a great amount of carbon dioxide dissolved.

Expt. 249.—Completely fill a flask with water and attach a cork and delivery tube, dipping under a jar of water inverted in a trough. Heat the flask and observe that dissolved air is driven out. Collect the air and see that, since it supports combustion, it contains oxygen. Prove also (by means of lime-water) that it contains carbon dioxide.

Hard and Soft Waters.—It is a fact familiar to every one that soap lathers very easily in some waters and not at all in others. If rain-water be used, the lathering takes place with great ease, while with the water which is supplied to some towns a lather can only be made with difficulty; and if we attempt the same process in sea-water there is no lathering at all. *Those waters in which soap lathers easily are said to be soft. When such is not the case the water is spoken of as hard.*

Temporary and Permanent Hardness.—Hard waters differ among themselves. Some can be softened by mere boiling, and when this is so the hardness is said to be *temporary*. If the hardness is not removed after the water has been boiled, and the water requires the addition of a chemical to soften it, such hardness is termed *permanent*.

> EXPT. 250.—Filter and then evaporate to dryness some samples of a water which you have found to be hard. Observe a white solid is left. Add to this white solid a little hydrochloric acid. Observe that it effervesces.
>
> EXPT. 251.—Filter and then evaporate samples of soft water, if any are obtainable. You will most probably find that very little residue remains and that no effervescence results on the addition of acid.

Hence the hardness of water is due to some white solid dissolved in the water, and this solid effervesces on the addition of acid.

We shall return to this in a later chapter.

CHIEF POINTS OF CHAPTER XVI.

Water is a clear liquid with a blue-green colour. It boils at 100° C., when it is converted into steam. It freezes at 0° C., becoming ice. Its density at 4° C. is 1. It possesses great solvent power.

Hydrogen is contained in Water. This has been found by causing heated iron to drive it out of the water, when the iron combines with the oxygen which is left, to form oxide of iron. Sodium turns the hydrogen out of water at ordinary temperatures. The hydrogen can be collected and examined.

Preparation of Hydrogen.—Hydrogen is best prepared by acting upon a dilute acid with a metal. Sulphuric acid and zinc were found to be suitable.

Sulphuric Acid and Zinc form Hydrogen and Zinc Sulphate.

Properties of Hydrogen. It is a colourless, odourless gas, considerably lighter than air. It does not support combustion, but itself burns in air. It forms an explosive mixture when mixed with air.

Production of Oxide of Hydrogen.—By collecting the product of combustion when hydrogen is burnt, and examining it, it is found to be a clear liquid with density 1, boiling point 100° C., freezing point 0° C. The product, or oxide of hydrogen, is thus seen to be really water. *Hydrogen in burning produces water which is, therefore, an oxide of hydrogen.*

Composition of Water by Volume.—This can be determined by means of an **Eudiometer**, into which known volumes of pure dry oxygen and hydrogen are bubbled after the eudiometer has been filled with mercury and inverted in a trough containing the same liquid. When an electric spark is passed through the mixture parts of the gases combine; and the experiment, when performed as described in the chapter, shows that *two volumes of hydrogen combine with one volume of oxygen to form water.* This process of building up a compound from its elements is called **synthesis.**

The opposite process of splitting a compound up into its elements is called **analysis**. The analysis of water is effected by passing an electric current through water contained in a voltameter. The result of such an experiment is precisely the same as that obtained, by synthesis, with the eudiometer.

Composition of Water by Weight.—Pure dry hydrogen is passed over heated oxide of copper. The hydrogen combines with the oxygen of the oxide, forming water, and leaves the metallic copper behind. The water formed is collected and weighed. The copper oxide is weighed before and after the experiment. Its loss in weight tells us the amount of oxygen in the water formed. The difference between the weight of the water formed and the oxygen it contains tells us the weight of the hydrogen in the water. When carefully performed, the experiment shows that *water is formed of eight-ninths its weight of oxygen and one-ninth its weight of hydrogen.*

Natural Waters generally contain dissolved materials. When the amount is very large the water is called a *mineral water*. Natural waters containing sulphuretted hydrogen are called *sulphur-waters*; those containing compounds of iron are called *chalybeate* waters; and those containing much carbon dioxide are spoken of as *effervescent.*

Hard and Soft Waters.—Those waters in which soap lathers easily are said to be soft. When such is not the case the water is spoken of as hard.

Hard waters which can be softened by boiling are said to have only *temporary* hardness. If the water cannot be thus softened, but requires the addition of a chemical, its hardness is said to be *permanent.*

QUESTIONS ON CHAPTER XVI.

1. What is the composition of water? Describe experiments by which its composition by volume may be found.

2. By what means is hydrogen most conveniently obtained? What are the principal properties of the gas?

3. Give diagrams and a short description of an experiment to prove that water is produced by the combustion of hydrogen.

4. Describe the preparation and collection of hydrogen by the action of (*a*) sodium, (*b*) iron, on water. What other products are also formed? How would you obtain each of the products? Describe briefly their various characters.

5. In the preparation of hydrogen from zinc and sulphuric acid, what other product is formed? How would you separate it from the liquid? Describe briefly its appearance.

6. Describe three experiments to show the most characteristic properties of hydrogen.

7. If twenty volumes of hydrogen and twenty-five volumes of oxygen be mixed in an eudiometer and exploded, what volume of gas would be left, and how would you show which gas it was?

8. What is the action of hydrogen on heated copper oxide? How can this action be used to find the composition of water by weight?

9. What tests would you apply in order to determine whether a given colourless, odourless, tasteless liquid was water or not?

10. What is meant by "hard" and "soft" water? How would you determine which of two samples of water was the harder?

11. In what respects do natural waters differ from pure water, and how are the differences brought about?

12. How could you prove that some water is produced during the combustion of coal gas?

13. What is an eudiometer? Give a sketch of one and explain for what purposes it may be used.

14. How would you prove that when hydrogen is passed over heated copper oxide something is taken away from the latter compound?

15. You are required to prove experimentally that in the previous question the part abstracted is really oxygen. How would you endeavour to do so?

16. Describe the apparatus you would use for the production and collection of hydrogen gas; name the materials required and describe the properties of the gas.

17. Write down what you consider to be the physical and chemical characteristics of water, that is, the properties which are possessed by water, but by no other substance.

18. Describe a method of liberating hydrogen from water (*a*) at a red heat; (*b*) at ordinary temperatures.

CHAPTER XVII

SULPHUR AND ITS COMPOUNDS

Physical Properties of Sulphur.—It will be well to first carefully examine the physical properties of sulphur. It is a brittle, yellow solid, which may be easily reduced to a fine powder. It is insoluble in water, but as we have already seen, it is soluble in carbon disulphide (Expt. 208). It also dissolves to a slight extent in turpentine. When heated it melts, and the melting point, *i.e.*, the temperature during melting, may be determined in the following manner :—

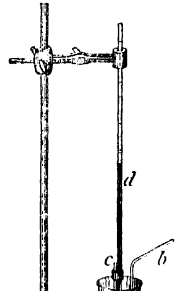

Fig. 127.—How to determine the Melting Point of Sulphur.

Expt. 252.—Draw out, in a blow-pipe flame, a piece of glass tubing so as to make a small thin-walled tube, about 2 or 3 inches long and ·1 inch in diameter. In this place some finely powdered sulphur and fix the tube to the bulb of the thermometer (it will probably stick to it if moistened ; if not, platinum wire is best). Place the thermometer bulb with the lower part of the tube in a beaker of sulphuric acid and heat it *gently*. Watch the sulphur carefully, and directly you see it to be melting note the temperature recorded by the thermometer, that is, the melting point of the sulphur. You will find it to be about 114° C.

Effects of Heat upon Sulphur.—The effect of raising the temperature above the melting point should be also observed, and the following experiments should be carefully performed.

EXPT. 253.—Put some finely powdered sulphur in a test-tube, using sufficient to fill the test-tube to a height of about 1½ to 2 inches, and heat very carefully in a small flame. When all has melted pour a little into cold water. Continue heating the remainder until the liquid boils, and again pour a little into cold water. Examine the cooled products.

If you performed this experiment with care and have observed the tube continuously, you should have found that the sulphur melts, at first, to a clear yellow mobile liquid, which when poured into cold water solidifies to the ordinary yellow sulphur. On further heating, however, the liquid gets much darker and thicker, becoming so viscid that the tube may even be inverted without the liquid running out; this occurs at a temperature of about 250° C. At higher temperatures still, it again becomes thin and mobile and finally boils, giving off a dark orange-red vapour which condenses to an orange liquid or to a yellow powder, or which burns at the mouth of the tube with a pale blue flame, and the characteristic smell of sulphur dioxide.

Varieties of Sulphur.—The boiling sulphur when suddenly cooled by cold water forms a remarkable product, as it may now be pulled about like a piece of caoutchouc, which indeed it closely resembles in appearance. It is called **Plastic Sulphur.**

EXPT. 254.—Place the piece of dry plastic sulphur, previously weighed, on one side and examine again after a few days. You will find that it has returned to the ordinary modification, but that its weight is unchanged.

Hence, plastic sulphur, although so different in its appearance and physical properties, really consists of nothing but sulphur.

On the surface of the water, into which you poured the boiling sulphur, you will probably have observed a quantity of a very fine yellow powder, and this consists of sulphur which has condensed directly from the gaseous to the solid state without forming the intermediate liquid, and is known as "flowers of sulphur." Sulphur is usually brought into commerce either in this form or as long solid rolls, both forms being obtained from a crude natural product which chiefly occurs, mixed with mineral impurities, in the volcanic regions of Italy and Sicily. The impure material is heated, and the melted sulphur is run off from the earthy impurities. It is then again melted and cast in cylindrical moulds to obtain the roll sulphur; or, it is strongly heated in retorts, so that it boils and the vapours are allowed to condense in a large cooled chamber, where they collect upon the walls and floor in the form of "flowers of sulphur."

Sulphur also occurs naturally in the combined state, the most important of its natural compounds being the sulphides, *i.e.*, compounds of sulphur with metals—and many of these are important *ores*, *i.e.*, are the natural compounds, from which we obtain the metal itself—for example, copper pyrites, galena or lead sulphide, blende or zinc sulphide, &c.

Expt. 255.—Heat a little iron pyrites in a test-tube. Observe melted sulphur on the walls of the tube, and a strong smell of sulphur dioxide.

Crystalline Sulphur.—Besides the two varieties of sulphur we have already observed, we may obtain sulphur in the crystalline form by two different methods.

Expt. 256.—Dissolve some sulphur in carbon disulphide and allow the solvent to slowly evaporate. Examine the crystals produced and sketch the most perfect.

The crystals so obtained are 8-sided, and belong to what is known as the *rhombic system*. The ordinary roll sulphur consists of minute crystals of this form.

Expt. 257.—Place some sulphur in an evaporating basin and heat carefully. When all has melted allow it to cool slowly, and as soon as a solid film has formed over the surface quickly pierce two holes in it, and through one of them pour out the interior liquid sulphur. Take up the top crust and examine the mass, which will be found to consist of a number of beautiful, needle-shaped, yellow, translucent crystals.

The crystals thus obtained are quite different from those produced in the previous experiment, and belong to what is known as the *monoclinic system*.

EXPT. 258.—Allow these crystals to remain in air for a few days, and examine again.

They will be now found to have become opaque, and have in fact changed to the ordinary form of sulphur, each crystal becoming converted to a number of minute rhombic crystals.

Allotropy.—We have now formed at least three modifications of sulphur, viz., the **rhombic**, **monoclinic**, and **plastic** sulphur, all consisting solely of sulphur. This property, which some elements possess, of existing in different forms is known as *allotropy*, and the several modifications are known as *allotropic forms*. A fourth allotropic form of sulphur also exists: a white powder insoluble in carbon disulphide, and which is present in flowers of sulphur.

Oxides of Sulphur.—We have already learnt that sulphur in burning combines with oxygen to form an oxide which we call sulphur dioxide, and which we know to be a colourless gas with a very pungent smell, soluble in water, forming an acid solution (p. 222). As, however, when so prepared it is always mixed with air or nitrogen, we must obtain it by some other method in order to have it in a purer state.

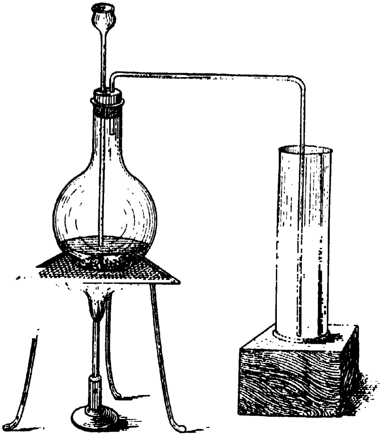

FIG. 128.—Preparation of Sulphur Dioxide.

Expt. 259.—Place a little copper in a test-tube with some strong sulphuric acid. There is no action. Heat the tube, and observe the smell of sulphur dioxide.

Expt. 260.—In a flask fitted with a cork and delivery tube place some copper turnings and some strong sulphuric acid, and heat over wire gauze by means of a Bunsen flame. As the gas is soluble in water, it cannot be collected in the same way as hydrogen and oxygen. Being heavier than air, however, it may be collected in the manner shown in the diagram (Fig. 128). Obtain by this means two jars of the gas.

In (1) place a lighted taper, and in (2) a piece of moist red flannel or a coloured flower, allowing it to remain for some time.

Properties of Sulphur Dioxide.—By these experiments you will have learnt that the gas does not burn, does not support combustion, and that it has the power of bleaching vegetable colours. This important property is largely made use of in technology, the gas being employed for bleaching, especially for delicate materials such as straw, silk, &c.

Sulphurous Acid and Sulphites.—Expt. 261.—Using the flask you have already fitted for the preparation of the gas, allow the sulphur dioxide to bubble through a solution of caustic soda, which, we have already seen, has the power of turning red litmus blue, and is called an alkali. Observe that the gas is absorbed, and that after a time the liquid smells of the gas and has become slightly acid. Boil the solution and observe that the smell disappears, showing that the gas is expelled by heating, while by further evaporation a white solid is ultimately left. Examine this solid, which is evidently soluble in water, and satisfy yourself that it is not caustic soda, having no action on litmus. Add to it a little sulphuric acid and observe the smell of sulphur dioxide, indicating that this was previously contained in the solid.

The formation of this solid is an example of the production of a **salt**, and we shall find later that many other salts are formed by analogous methods. The salts produced by neutralising **sulphurous acid**, as the solution of sulphur dioxide is called, are known as **sulphites**, and the particular salt we have thus prepared is *sodium sulphite*. All the sulphites are characterised by their property of giving off sulphur dioxide when acted upon by sulphuric or other strong acid.

EXPT. 262.—The flask which contained the copper and sulphuric acid should next be examined. It will be observed that most of the copper has disappeared, and the contents of the flask have a black colour. Add some water, shake well, and filter, when a blue liquid is obtained. Evaporate this partially, and put aside to crystallise. It will be found that blue crystals are so obtained.

These crystals are what are known commercially as *blue vitriol*, the chemical name being *sulphate of copper*. We may therefore state that **Copper and sulphuric acid produce sulphur dioxide and copper sulphate.**

Sulphur Trioxide and Sulphuric Acid.—Although sulphur dioxide does not burn, yet by suitable means we can make it combine with more oxygen and produce a higher oxide, which is known as **sulphur trioxide**.

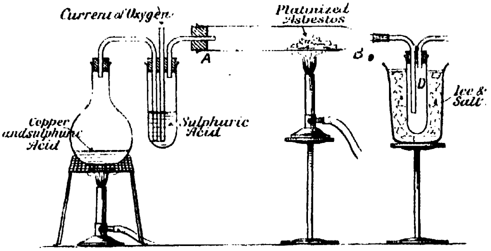

FIG. 129.—Preparation of Sulphur Trioxide.

EXPT. 263.—In the tube AB place some platinised asbestos,[1] and pass through the tube a well-dried mixture of sulphur dioxide and oxygen or air, the gases then being led into the cooled tube D. Heat the tube of asbestos, and note that white fumes are formed which condense in the cooler flask to form a white powder or crystals.

This powder is sulphur trioxide, and we may now state:—
Sulphur dioxide with oxygen forms sulphur trioxide.

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[1] This is prepared by dipping asbestos fibres into (1) platinic chloride solution, (2) ammonium chloride solution, and heating strongly. By this means the asbestos becomes coated with a quantity of very finely divided platinum.

EXPT. 264.—Add a little water to the trioxide in the flask, and test the liquid with litmus papers.

The sulphur trioxide dissolves very readily in water, evolving much heat and producing a hissing sound during solution. The solution so obtained is strongly acid, being, in fact, **sulphuric acid**.

EXPT. 265.—Heat the solution in an evaporating basin, and also, at the same time, in another evaporating basin, heat some dilute sulphuric acid, continuing the heating until thick white fumes are obtained, and satisfy yourself the liquid is sulphuric acid.

Sulphuric acid.—This very important compound, known also as *oil of vitriol*, must be examined, as it is one of the most important and useful of chemical products, being of the greatest service, not only in purely chemical operations, but also in various industrial and manufacturing processes.

It is a heavy, oily liquid, which, when strongly heated, boils at 335° C., and gives off a quantity of choking, pungent, white fumes, as has been observed in Expt. 265. It mixes with water in all proportions, and produces during the solution so much heat that the temperature may rise above 100° C., the boiling point of pure water, so that care has to be taken when sulphuric acid and water are mixed. It absorbs moisture from the air or from moist gases, and on this account is very frequently used for drying gases (Expt. 229), and, owing to the same affinity for water, it chars organic matter, such as wood, &c.

EXPT. 266.—In a little strong sulphuric acid place small pieces of wood, cloth, &c., and observe the effect.

Like most other strong acids, it burns the skin and destroys cloth, so that care must be always taken in its use.

Action of Sulphuric Acid on Metals.—Its action upon metals has been already seen: with some metals, *e.g.*, zinc, it reacts when cold and dilute, liberating hydrogen and forming a *sulphate* of the metal; with others, *e.g.*, copper, it has no action until heated, when it produces a *sulphate*, but with the liberation of *sulphur dioxide*.

We can understand the latter action if we remember that the sulphuric acid may be regarded as sulphur trioxide + water.

The hydrogen that we should have expected to be liberated has then abstracted some of the oxygen from this trioxide and so liberated the dioxide.

Sulphates.—With alkalis sulphuric acid also forms salts, just as was found in the case of sulphurous acid.

> EXPT. 267.—To a little sulphuric acid, in an evaporating dish, add drop by drop a solution of caustic soda, until the liquid has only a faint acid action. Evaporate partially and allow to crystallise and observe that colourless crystals are obtained.

These crystals are known as *sodium sulphate*, the salts obtained from sulphuric acid being called *sulphates* just as those from sulphurous acid were called *sulphites.* We have already prepared three sulphates, viz., sodium sulphate, zinc sulphate, and copper sulphate, the last of these being distinguishable by its fine blue colour.

Manufacture of Sulphuric Acid.—In the manufacture of sulphuric acid on the large scale, sulphur dioxide is first obtained, usually from iron pyrites (Expt. 255). The sulphur dioxide is then oxidised, but for this purpose a compound of nitrogen and oxygen is used—known as nitrogen peroxide—which readily gives up its oxygen to the sulphur dioxide, being itself converted to a compound of nitrogen and oxygen which contains less oxygen than the peroxide, and is known as nitric oxide. This nitric oxide, however, has the property of taking up oxygen from the air and becoming reconverted to the peroxide, which can again give up the oxygen to more sulphur dioxide, and thus acts as a carrier of oxygen from the air to the sulphur dioxide. Hence, in the presence of steam, sulphuric acid is formed. The sulphur dioxide, oxides of nitrogen (obtained from nitric acid), steam and air are passed into large leaden chambers where the reactions indicated above take place, with the production of sulphuric acid, which collects on the floor of the chambers and is afterwards concentrated by distilling off the water from glass or platinum retorts.[1]

Sulphides.—We have already found that copper and sulphur on being heated together unite to form a compound known as copper sulphide. This experiment should be again performed, using iron in place of copper. By this means we obtain a dark gray solid, *iron sulphide*, another example of a large class of compounds, which consist of a metal and sulphur united together, and known as the metallic sulphides.

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[1] It may be here stated that in reality the reaction is more complex than that indicated above, the formation of intermediate products having been neglected and only the final products considered.

> EXPT. 268.—Place a very small quantity of the iron sulphide in a test-tube and add a little hydrochloric or sulphuric acid. Note that it effervesces and that a gas is evolved.

Sulphuretted Hydrogen.—It has in this way been found that by the action of acid upon some metallic sulphides we obtain a gas with a peculiar odour, resembling that of rotten eggs. We must now collect and examine this gas.

> EXPT. 269.—Fit up a flask with cork, delivery tube, and thistle funnel, as for the preparation of hydrogen, and in the flask place some iron sulphide. Pour down the thistle funnel some hydrochloric acid diluted with about twice its volume of water (see that the thistle funnel dips below the liquid). Collect the gas over water in the usual way.
>
> EXPT. 270.—In a jar of the gas place a lighted taper. Note the extinction of the taper and the combustion of the gas with the smell of sulphur dioxide and also the formation of a pale yellow deposit on the inside of the jar.
>
> EXPT. 271.—Apply a light to the gas issuing from the delivery tube and over the flame hold a cooled vessel, *e.g.*, a flask of water. Observe that drops of liquid condense on the sides of the flask and that a yellow deposit is also formed. By collecting this on the end of a wire and burning it, satisfy yourself that it is sulphur.

The liquid may also be collected in the manner described in Expt. 244, and as in that case it will be here again found that its physical properties prove it to be water.

Hence the gas contains *hydrogen* (since water produced from it contains hydrogen, which could only have come from the gas, not being present in the air) and also *sulphur*, as shown by the formation of a sulphur deposit or of sulphur dioxide. It cannot be proved to contain any other element, and we may hence call it *hydrogen sulphide*, or, as it is generally termed, *sulphuretted hydrogen*. When the gas burns in a plentiful supply of air, sufficient oxygen is at hand to combine with both the hydrogen and sulphur, and hence sulphur dioxide is produced by the combustion ; but if there is an insufficient supply of oxygen, or if the flame is cooled, the hydrogen alone completely burns and the sulphur only partially, so that a quantity of sulphur is deposited uncombined.

EXPT. 272.—Pass the gas through solutions of (1) copper sulphate, (2) lead acetate, (3) an acid solution of arsenious oxide (N.B.—This is a *very strong poison*), and observe that in each case a thick precipitate is formed (1) black, (2) black, (3) yellow.

These precipitates consist of the *sulphides* of the metals, thus (1) is copper sulphide, (2) is lead sulphide, (3) is a sulphide of arsenic.

EXPT. 273.—Filter off the precipitate (1) and evaporate the solution, and hence observe that it consists of sulphuric acid.

Double Decomposition.—We thus find that from copper sulphate and sulphuretted hydrogen we obtain copper sulphide and sulphuric acid. As the sulphuretted hydrogen is hydrogen sulphide, it hence appears that the sulphuric acid is the hydrogen sulphate, that is, the copper sulphate is sulphuric acid with the hydrogen replaced by the copper. Such a change as this is called a *double decomposition*, both reacting compounds being converted into new products by the exchange of some constituents.

In a similar manner the lead acetate and hydrogen sulphide yield lead sulphide and hydrogen acetate (acetic acid).

EXPT. 274.—Pass some sulphuretted hydrogen into water, and observe that the water smells like the gas itself and behaves in the same way towards copper sulphate, &c.

The gas is therefore soluble in water, and its solution is frequently used in the chemical laboratory instead of the gas itself, being employed in qualitative analysis for the separation (or detection) of metals which, like copper, are precipitated as sulphide, from those which are not so precipitated.

EXPT. 275.—Allow this solution to stand a few days, and observe that it has become tinted and milky, and does not smell (or at any rate not so strongly) of the gas.

This change is due to the fact that the oxygen dissolved in the water slowly replaces the sulphur of the sulphuretted hydrogen and so forms water, while the displaced sulphur is precipitated in the form of a white powder.

Quantitative Composition of Sulphuretted Hydrogen.—EXPT. 276.—In order to determine, at least partially, the composition of sulphuretted hydrogen, a tube (Fig. 130) should be about half filled with the gas over mercury (since it is slightly soluble in water). In the bent portion a piece of tin is placed, and strongly heated by a Bunsen burner, the end of the tube being closed to prevent the mercury being driven out by the expansion of the gas. The tin, when thus heated, abstracts the sulphur, and forms tin sulphide, leaving the hydrogen. Observe the volume of the hydrogen when the tube is again cool and see that the gas is really hydrogen.

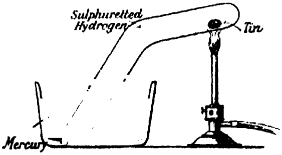

FIG. 130.—Composition of Sulphuretted Hydrogen.

It is thus found that **Sulphuretted hydrogen contains its own volume of hydrogen.**

For the above experiment and others in which a solid has to be heated in a closed volume of a gas, the apparatus shown in Fig. 131, designed by Mr. C. M. Stuart, Head Master of St. Dunstan's College, Catford, is very convenient. The U-tube contains a small quantity of mercury, just sufficient to fill the bend, and the solid is placed in the bulb of the tube *ca*, which is then fixed to the U-tube by indiarubber tubing as shown. The gas is passed through from *b* until it is considered that all the air is displaced. A little more mercury is then poured into the U-tube, the level *xy* noted, and the drawn-out end *a* having been sealed by a Bunsen flame the solid may be heated.

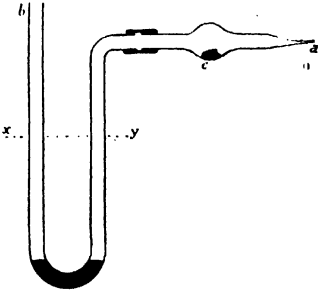

FIG. 131.—Arrangement for Heating a Solid in a closed Volume of Gas.

Chief Points of Chapter XVII.

Physical Properties of Sulphur.—It is a brittle, yellow solid, which may easily be reduced to a fine powder. It is insoluble in water, but dissolves in carbon disulphide, and also to a small extent in turpentine. It melts at about 114° C. to a clear, yellow, mobile liquid, which when poured into cold water solidifies to ordinary yellow sulphur. On further heating the yellow liquid becomes darker in colour, and more viscid, until at about 250° C. it will not run out, even though the vessel containing it is inverted. At still higher temperatures the liquid again becomes thin and mobile and finally boils, evolving a dark orange-red vapour, which condenses either to an orange liquid or to a yellow powder. If the boiling sulphur be poured into cold water it solidifies to a solid resembling caoutchouc.

Varieties of Sulphur.—This elastic solid is called *plastic sulphur*. If left in contact with air it returns to ordinary sulphur in a few days without any change of weight. The yellow powder into which sulphur vapour condenses, without passing through an intermediate liquid state, is called *flowers of sulphur*. Ordinary commercial sulphur is called *roll sulphur*.

Crystalline Sulphur.—The crystals left when a solution of sulphur in carbon disulphide is allowed to evaporate belong to what is called the *rhombic system*. Those obtained from melted sulphur in the manner described in the chapter belong to the *monoclinic system*. The monoclinic sulphur crystals will, if left alone, gradually change back to the rhombic variety.

Allotropy is the property some elements, like sulphur, possess of existing in different forms which are known as *allotropic forms*. The allotropic forms of sulphur are four in number, viz., rhombic, monoclinic, plastic, and amorphous.

Oxides of Sulphur.—*Sulphur dioxide* is formed when sulphur burns in air or oxygen. It is also given off when copper is heated with strong sulphuric acid. It is a gas with a pungent smell which does not burn nor support combustion and has the power of bleaching vegetable colours. It dissolves in water to form *sulphurous acid*. When sulphurous acid is neutralised by alkalis, *salts* called *sulphites* are obtained.

Sulphur Trioxide.—By suitable means sulphur dioxide can be made to combine with more oxygen to form a higher oxide known as sulphur trioxide. This oxide dissolves in water, with a hissing noise accompanied by the evolution of much heat, to form sulphuric acid.

Sulphuric Acid is a heavy, oily liquid, which boils at 335° C., giving off choking, pungent, white fumes. It mixes with water in all proportions with the evolution of much heat. It absorbs moisture very readily and is consequently used for drying gases. For the same reason it chars any organic substance it comes in contact with. Sulphuric acid forms salts called *sulphates*.

Manufacture of Sulphuric Acid.—This depends upon the fact that Sulphur dioxide can be made to unite with more oxygen to form sulphur trioxide, which dissolves in water, to form sulphuric acid. Nitrogen peroxide, which is obtained from nitric acid, easily gives up the oxygen required to combine with the sulphur dioxide. The nitrogen peroxide is thus converted into nitric oxide, which, by combining with oxygen from the air, is again changed to nitrogen peroxide, and all the changes are repeated again and again. In practice, the sulphur dioxide, oxides of nitrogen, water in form of steam, and air, are all passed into large leaden chambers on the floor of which the sulphuric acid, obtained by the changes we have described, collects.

Sulphides are compounds of sulphur with some other element. If sulphur is combined with a metal a *metallic sulphide* is the result, such as zinc sulphide or blende.

Sulphuretted Hydrogen, or hydrogen sulphide, is obtained by the action of dilute acids upon certain sulphides. It is a gas with an odour of rotten eggs. It burns in air with a pale blue flame, forming sulphur dioxide and water. If, however, there is only a small supply of air or oxygen, water is formed and the sulphur deposited as a yellow powder. Sulphuretted hydrogen will not support combustion. When the gas is passed into solutions of certain metallic salts, precipitates consisting of metallic sulphides are thrown down.

Double Decomposition.—Those changes where both reacting compounds are converted into new products by the exchange of some constituents, are known as double decompositions, thus:—

Copper sulphate and sulphuretted hydrogen give Copper sulphide and sulphuric acid.

Quantitative Composition of Sulphuretted Hydrogen.—By heating a piece of tin in sulphuretted hydrogen, contained in a bent tube over mercury, the tin is made to combine with the sulphur and to leave the hydrogen. If when the apparatus is cool the volume of the hydrogen is measured, it is found to be the same as that of the sulphuretted hydrogen taken at first.

QUESTIONS ON CHAPTER XVII.

1. Describe the changes which sulphur undergoes when heated.
2. What is plastic sulphur, and how is it obtained? How would you prove it consists solely of sulphur?
3. Describe two methods for obtaining crystals of sulphur. What differences are there in the crystals so obtained?
4. What happens when sulphur burns? By what other method can you obtain the product formed?
5. Give an account of the properties of sulphur dioxide.
6. What are sulphites, and how are they obtained? By what characteristic reaction would you recognise a sulphite?
7. Describe, with a diagram, an experiment to prove that when sulphur burns the product formed weighs more than the original sulphur.
8. Describe the appearances and properties of sulphur trioxide. How is it obtained, and what is its action on water?
9. Give a short account of the method of preparation of sulphuric acid.

10. What are sulphates? Describe a method of preparation, and give a *short* account of the appearance and properties of three sulphates.

11. How may sulphide of copper be obtained? What differences are there in the chemical nature, appearance, and properties of sulphide of copper and sulphate of copper?

12. If iron and sulphur be heated and the product so obtained treated with hydrochloric acid, what gas is obtained? Give an account of its more characteristic properties.

13. How could you prove that sulphuretted hydrogen contains sulphur and hydrogen?

14. Sulphuretted hydrogen is passed through solutions of (1) copper sulphate, (2) lead acetate. Describe what happens, and explain the nature of the chemical changes.

15. If in preparing hydrogen from iron and hydrochloric acid there happened to be sulphide of iron mixed with the iron, with what would the hydrogen be mixed, and how could you obtain this impurity in a pure state?

16. What are "flowers" of sulphur and how is sulphur obtained in this condition?

17. By what means may it be proved that sulphuretted hydrogen contains its own volume of hydrogen?

18. What is meant by a double decomposition? Give three examples of such changes.

19. What difference is there in the chemical behaviour of a sulphite and a sulphate? How would you detect the presence of sodium sulphite in a quantity of sodium sulphate?

20. What do we mean by allotropic forms? Give examples.

CHAPTER XVIII

SALT

Properties of Salt.—We shall now investigate common or table salt, and the various products which we can obtain from it. At the outset salt itself should be carefully examined, and a table prepared of its characteristic properties. It will be thus found to be a white soluble powder, which forms crystals in the form of small cubes, containing no water of crystallisation. It does not suffer any apparent change unless heated very strongly, when it can be fused or melted. If a little be placed in the non-luminous flame of a Bunsen burner the flame becomes coloured bright golden-yellow.

Expt. 277.—Place a little salt in a test-tube, and add some strong sulphuric acid; warm slightly. Observe that a gas is evolved which possesses a powerful pungent odour, and forms white fumes in the air.

We must now endeavour to collect and examine this "salt gas."

Salt Gas.—Expt. 278.—Fit up a glass flask for the preparation of this gas from the above materials, and collect it as shown in the diagram (Fig. 128). When fumes form at the top of the jar in which the gas is being collected, the jar may be regarded as full. Fill two jars, and keep them closed until required by pieces of greased glass.

(1) Insert a lighted taper and observe that the gas does not burn, neither does it support combustion.

(2) Open under water, and observe that the water rushes up and completely fills the jar. (Probably owing to the

incomplete displacement of the air the gas will *not completely* fill the jar, but it does so if unmixed with air.) Test the solution with litmus papers. It is found to be *acid*.

We therefore find that the "salt gas" **is soluble in water, and forms an acid solution.**

EXPT. 279.—Pass the gas through a solution of caustic soda until the liquid has become slightly acid. Evaporate the solution to dryness, and observe that a white solid is left. Taste this solid, and satisfy yourself that it is salt.

EXPT. 280.—Do a similar experiment, using caustic potash solution in place of the solution of caustic soda. Observe that, as before, you obtain a white solid, which has a taste somewhat resembling that of salt. Heat this solid with a little sulphuric acid, and observe that the "salt gas" is again produced.

Composition of Salt.—Salt, therefore, results from the interaction of the "salt gas" with caustic soda. The solution of the "salt gas" is, in fact, the acid which is so frequently used under the name of *hydrochloric acid*, and the last experiment should be again performed with hydrochloric acid in place of the "salt gas" in order to verify this statement. If the contents of the flask or test-tube used in the preparation of the gas be examined and allowed to crystallise, crystals are obtained of a solid known as sodium sulphate, and we may state that **Common salt and sulphuric acid yield hydrochloric acid gas and sodium sulphate.**

Composition of "Salt" or Hydrochloric Acid Gas.—EXPT. 281.—In the flask A (Fig. 132) place some salt and strong sulphuric acid. Pass the "salt gas" so obtained over heated copper oxide in the hard glass tube BC. Observe that **water** collects in the test-tube D, and that the copper oxide is converted into a green substance. Examine this green substance and see that it also gives off the "salt gas" when acted upon by sulphuric acid.

As water is formed the "salt gas" must evidently contain hydrogen, and the simplest explanation of the experiment is that the hydrogen of "salt gas" combined with the oxygen of the copper oxide, forming the water, while the copper with the other part of the "salt gas" formed the green substance.

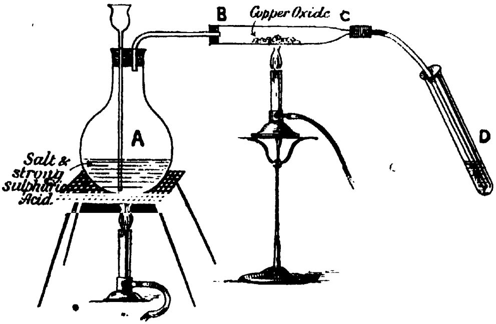

FIG. 132.—Composition of Hydrochloric Acid Gas.

EXPT. 282.—To further investigate this acid gas collect a tube full over mercury, and quickly introduce into it a piece of clean sodium. Allow it to stand, and observe that the volume of the gas becomes ultimately reduced to one-half the original volume (correction being made for the difference in pressure), while further, the sodium gets covered with a white powder which you may satisfy yourself is salt.

Test the gas left with a lighted taper and see that it has the properties of hydrogen.

We have thus found that from hydrochloric acid gas, sodium abstracts something and leaves one-half the original volume of hydrogen. Hence, also, salt consists of sodium with *something* and this same material with hydrogen forms hydrochloric acid gas. The name given to this material is **Chlorine,** and we shall now endeavour to obtain this element from the hydrochloric acid gas.

Preparation and Properties of Chlorine.—EXPT. 283.—In a fairly large flask (Fig. 133) put some manganese dioxide (a black powder which also occurs naturally in compact masses known as *pyrolusite*). Through the thistle funnel pour in sufficient hydrochloric acid (*i.e.* the strong acid used in the laboratory) to cover the oxide and see that the end of the thistle funnel is below the level of the acid. Observe that a dark greenish-brown liquid results. Warm and observe the formation of a heavy green gas which may be collected in the same manner as hydrochloric gas.[1] In this manner fill five jars, and close each jar tightly.

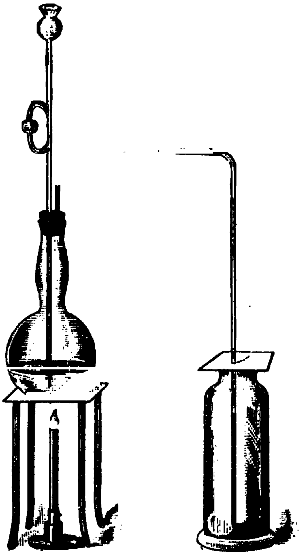

Fig. 133.—Preparation of Chlorine.

(1) In the first jar place some moist coloured rags, or flowers, and observe the bleaching. Writing in ordinary ink may also be bleached.

(2) Plunge in a lighted taper and observe that the taper burns, but with the formation of clouds of soot, and observe that water shaken up in the jar becomes acid.

(3) Place in some filter-paper moistened with turpentine and compare the result with that obtained in the previous experiment.

(4) Place in the jar, by means of a deflagrating spoon, some burning phosphorus, and observe it continues to burn with the formation of white fumes.

(5) Place in the gas some burning sodium, and see that it also burns with the formation of white fumes. Dissolve these fumes in water, and see that a solution of salt results.

Pass the gas through water for a short time, and observe that it is slightly soluble, the solution, known as chlorine water, smelling of the gas.

It is thus seen that sodium with this gas forms salt, and hence this gas is actually the second constituent of the hydrochloric acid gas, *i.e.*, *chlorine.* In the case of the burning taper the effect is due to the fact that the taper consists of two constituents, viz., carbon and hydrogen, and that, whereas both burn in air to form oxides, only one, the hydrogen, can combine with the chlorine, forming hydrochloric acid gas (hence the acidity of the water), while the carbon, being incapable of combining with the chlorine, gets deposited as soot. In the case of the burning phosphorus the white fumes consist of the compound of phosphorus and chlorine, *i.e.*, a *chloride of phosphorus*, and, similarly, hydrochloric acid gas may be called *hydrogen chloride*, and common salt *sodium chloride.*

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[1] The gas is best collected in the open air or in a draught cupboard.

We have seen above that hydrogen can combine with chlorine; indeed, so great is their tendency to unite that a mixture of hydrogen and chlorine explodes even if exposed to a bright light—as direct sunlight or burning magnesium—and combines gradually in diffused daylight, the mixture being seen to lose the green colour imparted to it by the chlorine. If the combination is effected in the eudiometer described on page 240, it is found that 1 volume of hydrogen combines with 1 volume of chlorine to form 2 volumes of hydrochloric acid gas, so that in this case there is no contraction during the formation of the compound. Also, if a concentrated solution of hydrochloric acid be decomposed by the electric current in the voltameter, equal volumes of hydrogen and chlorine result, although at the commencement of the experiment more hydrogen seems to be obtained, this being due to the solution of the chlorine.

Summary of Results.—We may summarise our results thus: Chlorine obtained from hydrochloric acid and manganese dioxide is a very heavy, yellowish-green gas, with an irritating, choking smell. It is soluble in water, the solution having an odour similar to that of the gas. It possesses powerful bleaching properties. Though it does not burn itself, it is capable of supporting the combustion of many bodies, with the formation of chlorides. Of these, sodium chloride is ordinary table salt, and hydrogen chloride is hydrochloric acid gas, which consists of equal volumes of hydrogen and chlorine united without contraction, this union occurring directly in daylight. The weight of a litre of chlorine [which may be found by weighing a flask full of (1) air or hydrogen, (2) chlorine, and allowing for the weight of air or hydrogen in (1)] is about 3·18 grams, so that it is about 35·5 times as heavy as hydrogen.

As when salt or other chloride is heated with sulphuric acid it gives rise to hydrochloric acid, it follows that chlorine may be obtained by heating a *chloride* with sulphuric acid and manganese dioxide, and this also forms a convenient test for *chlorides*.

Various Compounds of Chlorine.—Expt. 284.—Pass chlorine for some time through a *boiling* solution of caustic potash. Allow the solution to crystallise, and examine the crystals which first separate out. They will be found to consist of *potassium chlorate*, which evolves oxygen when heated, either alone or with a little manganese dioxide.

We thus obtain *potassium chlorate* from chlorine and hot potash, while some potassium chloride is also formed. The *chlorates* may be regarded as derived from an acid, which we may call *chloric acid*, just as the chlorides are derived from hydrochloric acid, and which differs from the latter acid in containing oxygen as well as hydrogen and chlorine.

Of the chlorates, *potassium chlorate* is by far the most important, being largely employed in the manufacture of matches and in pyrotechny.

Expt. 285.—Pass chlorine through a *cold* solution of caustic potash, and observe that you no longer obtain potassium chlorate.

In this case there results, as before, potassium chloride; the second product is, however, not potassium chlorate, but a compound containing a smaller percentage of oxygen, and termed *potassium hypochlorite*, which, by the action of acid, liberates chlorine. By employing lime in place of caustic potash a somewhat similar mixture results, consisting of lime with a compound of the chloride and hypochlorite. This also evolves chlorine when acted upon by an acid, and is hence very largely employed for bleaching, under the name of *bleaching powder*, which is manufactured on the large scale by the method indicated above.

CHIEF POINTS OF CHAPTER XVIII.

Common Salt is a white soluble powder which forms crystals in the form of small cubes, containing no water of crystallisation. When very strongly heated it melts. It imparts a golden-yellow colour to the flame. When heated in a tube with strong sulphuric acid, a gas with a powerful pungent odour and which forms fumes in air is given off. This gas we provisionally called "*salt gas.*"

Properties of "Salt Gas."—It neither burns nor supports combustion. It is very soluble in water forming an *acid solution*. If passed into a solution of caustic soda and the resulting solution evaporated to dryness a white solid is obtained which is found to be salt. Since the same results follow if this experiment is repeated, substituting hydrochloric acid for "salt gas," we are justified in calling "salt gas" by the name of hydrochloric acid gas. We may therefore state that *Common salt and sulphuric acid yield hydrochloric acid gas and sodium sulphate.*

Composition of "Salt-" or Hydrochloric Acid-Gas.—When the acid gas is passed over heated copper oxide, water is formed and can be collected. Water contains hydrogen, therefore probably this hydrogen was obtained from the "salt gas." The copper oxide becomes changed into a green substance; probably the copper of the oxide has combined with the other part of the "salt gas" to form the green compound.

The "salt gas" is *proved* to contain hydrogen, to the extent of one-half its volume, by the action of a piece of sodium upon a quantity of the gas in a tube over mercury. Half the amount of hydrogen is found left in the tube. The other constituent of the "salt gas" combines with the sodium to form salt again. The name given to this other constituent is *chlorine*.

By means of the eudiometer and the voltameter (*see* Chap. XVI.), the hydrochloric acid gas can be proved to be made up of one volume of hydrogen and one volume of chlorine combined to form two volumes of hydrochloric acid gas.

Preparation and Properties of Chlorine.—Chlorine is obtained from hydrochloric acid by heating it with manganese dioxide. The gas being heavier than air is collected by downward displacement as in Expt. 283.

Chlorine is a very heavy, yellowish-green gas with an irritating, choking odour. It is soluble in water and the solution smells like the gas. It possesses powerful bleaching properties. Though it does not itself burn, it is capable of supporting the combustion of many bodies, with the formation of *chlorides*. The weight of a litre of chlorine is about 3·18 grams, it is hence about 35·5 times as heavy as hydrogen.

Various Compounds of Chlorine.—*Potassium chlorate* can be obtained by passing chlorine for some time through a *boiling* solution of caustic potash and allowing the solution to crystallise. The *chlorates* are salts of an acid known as *chloric acid*. Potassium chlorate is the most important chlorate; it is used in the manufacture of matches and fireworks.

Potassium hypochlorite is obtained by passing chlorine through a *cold* solution of caustic potash. If lime is used instead of caustic potash, a mixture called *bleaching-powder* is obtained.

QUESTIONS ON CHAPTER XVIII.

1. How is hydrochloric acid obtained? Give a short account of its chief properties.

2. What are chlorides? How may they be obtained? Give examples.

3. Briefly indicate the reasoning which leads to the supposition that hydrochloric acid gas contains hydrogen united with another gas, and state how this second gas may be obtained from the acid.

4. What is the general effect of hydrochloric acid upon (*a*) metals, (*b*) oxides?

5. How may it be proved that hydrochloric acid gas consists of one half its volume of (*a*) hydrogen and (*b*) chlorine?

6. Describe the properties of chlorine, and state how you would obtain the gas from salt and then reconvert it into salt.

7. What are chlorates? How do they differ from chlorides?

8. Under what conditions does chlorine unite with (*a*) hydrogen, (*b*) phosphorus, (*c*) sodium?

9. A lighted taper is placed in a jar of chlorine, what happens, and why?

10. How may chlorine be (*a*) obtained from, (*b*) converted into hydrochloric acid?

11. What is the action of sulphuric acid upon salt? What are the properties of both *products*?

12. From hydrochloric acid how could you obtain (*a*) hydrogen, (*b*) common salt, (*c*) potassium chlorate?

13. What is bleaching powder? How is it obtained, and on what does its chief use depend?

14. By what chemical tests would you recognise a metallic chloride?

CHAPTER XIX

CHEMICAL NOMENCLATURE

Atoms and Molecules.—We are now in a position to understand the nomenclature employed by chemists to represent the various chemical changes which we have been studying. All matter is supposed to be built up of minute particles termed *atoms*. These usually exist combined with other atoms to form *molecules*, and these molecules are the smallest particles of matter which can have a separate existence. In a compound the molecules are composed of a number of different kinds of atoms united together. In an element the atoms are all of the same kind.

Chemists represent the atoms of the various elements by letters. These are in general the first letter or, sometimes, the characteristic letters of the name (or Latin name) of the element. Thus the letter H represents the atom of hydrogen, and O the atom of oxygen. To represent a compound, a **formula** consisting of a combination of symbols is employed. Thus water must evidently be represented by some combination of the letters H and O, but we do not yet know the number of H's and O's which we must employ.

Avogadro's Law.—In order to deduce the formulæ applicable to the different compounds, we must make use of a law first enunciated by the Italian chemist *Avogadro*, and which is now almost universally accepted by chemists. This law states that **Equal volumes of all gases under similar conditions of temperature and pressure contain equal numbers of molecules.**

Formula for Hydrochloric Acid Gas.—From what we have learnt concerning the formation of hydrochloric acid gas, we know that 1 volume of hydrogen combines with 1 volume of chlorine to form 2 volumes of hydrochloric acid gas, that is, by Avogadro's Law, a certain number of molecules of hydrogen combine with the same number of molecules of chlorine to form twice that number of molecules of hydrochloric acid gas; so that 1 molecule of hydrogen must combine with 1 molecule of chlorine to form 2 molecules of the compound. Moreover, as each molecule of hydrochloric acid gas must contain *at least* 1 *atom* of hydrogen and 1 *atom* of chlorine, then the 2 molecules must contain *at least* 2 atoms of each element, and therefore 1 molecule of hydrogen (which gives rise to 2 molecules of hydrochloric acid gas) must contain at least 2 atoms of hydrogen, and similarly, each molecule of chlorine must contain at least 2 atoms of chlorine. This is the number of the atoms in the molecule of these elements if the compound we have considered contains 1 atom of each element, and this, the simplest explanation, is that which really holds. The formula for hydrochloric acid gas may therefore be written as HCl. This result which we have found for hydrogen and chlorine, is also generally true for other gaseous elements; so that we may state, as a fairly general rule, that **The molecule of a gaseous element contains 2 atoms.**

Formula for Water.—The formation of water gives a slightly more complicated case, since it was found that 2 volumes of hydrogen with 1 volume of oxygen form only 2 volumes of water vapour. From this it follows, however, by Avogadro's Law, that 2 molecules of hydrogen with 1 molecule of oxygen form 2 molecules of water, and hence from what has been previously stated, 4 *atoms* of hydrogen combine with 2 *atoms* of oxygen to form 2 molecules of water. Each molecule of water therefore consists of 2 atoms of hydrogen with 1 atom of oxygen. We can consequently represent the molecule of water by the formula H_2O.

Chemical Equations.—We may now learn how to represent the formation of compounds by what are termed *chemical equations.* The formulæ of the reacting bodies are placed on the left hand side of the equality sign, and those of the products on the right.

(1) $$2H_2+O_2=2H_2O\,;$$ [1]

that is,

2 mols. hydrogen with 1 mol. oxygen form 2 mols. of water.

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[1] In this formula for water the small 2 only refers to the H, the large 2 to the whole group, and similarly in the other formulæ employed.

(2) $$H_2+Cl_2=2HCl.$$

1 mol. hydrogen with 1 mol. chlorine form 2 mols. hydrochloric acid gas.

Atomic Weights.—We have, however, found that any volume of oxygen is 16 times as heavy as the same volume of hydrogen, so that each molecule of oxygen is 16 times as heavy as a molecule of hydrogen. Consequently the ratio of the weight of the atoms must also be 16. So that if we regard the weight of an atom of hydrogen as our unit, the *atomic weight* of oxygen must be 16. **The atomic weight of an element is, therefore, the ratio of the weight of its atom to the weight of the atom of hydrogen.**

By similar reasoning the value 35·5 may be obtained as the atomic weight of chlorine (*see* page 266), and hence the symbols Cl, O, and H will represent respectively 35·5, 16 and 1 parts by weight of chlorine, oxygen, and hydrogen. A list of the symbols of the elements with their atomic weights is given on page 275, but we shall in general use simplified numbers, as 16 in place of (15·87) for oxygen.

We may, therefore, extend the meaning of the chemical equations given above and read thus:—

$$2H_2+O_2=2H_2O.$$

4 ($=2\times2$) parts by weight of hydrogen with 32 ($=2\times16$) parts by weight of oxygen unite to form 36 ($=\overline{2\times2}+\overline{2\times16}$) parts by weight of water.

Formulæ for Sulphur Dioxide and Sulphuretted Hydrogen.—It may be easily proved experimentally that when sulphur burns in oxygen there is no alteration of volume. From this, by an application of Avogrado's Law, it follows that from a certain number of molecules of oxygen we obtain the *same* number of molecules of sulphur dioxide. Hence each molecule of this compound is obtained from 1 molecule, *i.e.*, 2 *atoms*, of oxygen. This does not show how many atoms of sulphur are present, but if we assume the presence of 1 atom the formula would be SO_2. As we have previously found that sulphur gains 100 per cent. in weight on burning, it is clear that the 1 atom of sulphur has the same weight as 2 atoms of oxygen, *i.e.*, the atomic weight is $2\times16=32$. This number agrees with that obtained for the atomic weight of sulphur by other methods, and hence indicates that the formula adopted is correct, *i.e.*, sulphur dioxide may be represented by the formula SO_2. Sulphuretted hydrogen is also found to contain its own volume of hydrogen, and by analogous reasoning we obtain SH_2 for its formula.

Equivalent Weights and Valency.—The methods employed for finding the atomic weights differ in different cases. A method of frequent application is to find the weight of an element which can displace 1 gram of hydrogen from a compound, or which unites with 8 grams of oxygen, or with 35·5 grams of chlorine (these being the weights which unite with 1 gram of hydrogen). For example, it may be readily found that 12 grams of magnesium displace 1 gram of hydrogen from an acid,[1] or that 23 grams of sodium displace 1 gram of hydrogen from water. These values 12 and 23 may be called the **equivalent weights** of magnesium and sodium. It is, however, considered that while 1 atom of sodium replaces 1 atom of hydrogen, yet the atom of magnesium is capable of replacing 2 atoms of hydrogen. The atomic weight of sodium is therefore 23, but in the case of magnesium the atomic weight is 12 × 2, for we have found that the magnesium is 12 times heavier than the hydrogen it displaces, that is, each atom is 12 times heavier than 2 atoms of hydrogen, *i.e.*, 24 times the weight of 1 atom. The number of hydrogen atoms which 1 atom of an element can replace is called its **valency**, and from what has been said it will be seen that the **atomic weight = equivalent weight × valency**. It would, however, take us too far to discuss the methods of determining the valency, or other methods of finding the atomic weights.

Molecular Weights.—It follows from Avogadro's Law that the ratio of the weight of any volume of a gas to the weight of the same volume of hydrogen, is equal to the ratio of the weight of each molecule of the gas to the weight of the molecule of hydrogen; that is, the *density* of the gas referred to hydrogen is the ratio of the molecular weight of the gas to the molecular weight of hydrogen. Since the molecule of hydrogen contains 2 atoms, its molecular weight is 2; hence

$$\text{Density referred to hydrogen} = \frac{\text{Mol. wt. of gas}}{\text{Mol. wt. of hydrogen}} = \frac{\text{Mol. wt. of gas}}{2}$$

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[1] This experiment may easily be performed with the apparatus described in Expt. 296.

so that the **molecular weight of a gas is double its density** (referred to hydrogen).

Chemical Actions Represented by Equations.—We will now give a number of equations to represent some of the chemical reactions we have studied. In all the cases the quantitative relation should be found and written down. The decomposition of potassium chlorate by heat is represented by :—

$2KClO_3 = 2KCl + 3O_2$, or

$2 \times 122 \cdot 5$ $(= 39 + 35 \cdot 5 + 48)$ parts by weight of potassium chlorate	will yield	$2 \times 74 \cdot 5$ $(= 35 \cdot 5 + 39)$ parts by weight of potassium chloride

with 3×32 parts by weight of oxygen.

The burning of sodium, carbon, phosphorus, and sulphur are respectively represented by :—

$$2Na_2 + O_2 = 2Na_2O.$$
$$C + O_2 = CO_2$$
$$S + O_2 = SO_2.$$
$$P_4 + 5O_2 = 2P_2O_5.$$

(Write these also in words and, using the atomic weights on page 275, give the weights of the substances taking part in the reactions as is done in the above example.)

The displacement of hydrogen from water or acids, is given by the equations :—

$$2H_2O + Na_2 = 2NaOH + H_2.$$
$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$

It will be noticed that according to the first equation, only one half the hydrogen is eliminated from water by sodium, and it may be shown that caustic soda still contains hydrogen.

The formation of sulphur dioxide is more complicated :—

$$Cu + 2H_2SO_4 = CuSO_4 + SO_2 + 2H_2O\ ;$$

that is, copper forms with sulphuric acid copper sulphate, sulphur dioxide, and water ; while the formation of sulphur trioxide and sulphuric acid are very simply shown :—

(1) $2SO_2 + O_2 = 2SO_3.$

(2) $SO_3 + H_2O = H_2SO_4.$

The sulphide of iron we have used is represented by FeS (the symbol Fe standing for iron, *ferrum*), and the reaction of the sulphide with sulphuric acid by the equation :—

$$FeS + H_2SO_4 = FeSO_4 + SH_2.$$

The equation for the preparation of chlorine is frequently written in two stages :—

(1) $MnO_2 + 4HCl = MnCl_4 + 2H_2O$

(2) $MnCl_4 = MnCl_2 + Cl_2$.

This indicates that an intermediate compound ($MnCl_4$) is formed, which, however, decomposes to form a lower chloride

TABLE OF ELEMENTS WITH THEIR SYMBOLS AND ATOMIC WEIGHTS.

Element.	Symbol.	At. Wt.	Element.	Symbol.	At. Wt.
Aluminium .	Al	27·02	Mercury	Hg	200
Antimony ...	Sb	119·6	Molybdenum.	Mo	95·7
Argon	A	39·8(?)	Nickel	Ni	58·6
Arsenic.........	As	74·9	Niobium ...	Nb	94
Barium... . ..	Ba	136·8	Nitrogen	N	14·01
Beryllium ...	Be	9·1	Osmium	Os	190·3
Bismuth	Bi	207·5	Oxygen.........	O	15·87
Boron	B	10·9	Palladium ...	Pd	106·2
Bromine	Br	79·7	Phosphorus ..	P	31
Cadmium	Cd	112	Platinum ..	Pt	194·3
Cæsium.........	Cs	132·7	Potassium .	K	39
Calcium	Ca	39·9	Rhodium	Rh	103
Carbon	C	11·97	Rubidium ..	Rb	85·2
Cerium	Ce	140	Ruthenium ...	Ru	102
Chlorine	Cl	35·3	Samarium ...	Sm	150
Chromium ..	Cr	52·4	Scandium... ..	Sc	44
Cobalt	Co	58·6	Selenium	Se	79
Copper	Cu	63·2	Silicon	Si	28·3
Didymium ..	Di	143	Silver	Ag	107·9
Erbium.........	Er	166	Sodium.........	Na	23
Fluorine	F	19	Strontium ...	Sr	87·5
Gallium	Ga	69·8	Sulphur	S	32
Germanium ...	Ge	72·3	Tantalum.. ..	Ta	183
Gold	Au	196·7	Tellurium .	Te	125
Helium.........	He	4·4	Thallium	Tl	204
Hydrogen ...	H	1	Thorium ... ,	Th	232
Indium.........	In	113·4	Tin	Sn	117·7
Iodine	I	126·6	Titanium	Ti	48
Iridium..... ...	Ir	192·6	Tungsten	W	184
Iron	Fe	55·9	Uranium	U	239
Lanthanum ...	La	138	Vanadium ..	V	51·1
Lead	Pb	206·3	Ytterbium .	Yb	173
Lithium	Li	7	Yttrium	Y	89
Magnesium .	Mg	24	Zinc	Zn	65
Manganese ...	Mn	54·8	Zirconium ...	Zr	90·4

of manganese and chlorine. As, however, it is improbable that the compound $MnCl_4$ exists, it is preferable to write the reaction in one equation :—

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2.$$

The student should endeavour to represent other actions, as the burning of sulphuretted hydrogen, &c., by equations, taking great care that the same number of atoms of each element occur on the two sides of the equation, and paying attention to the known formulæ and the valencies given. It must be remembered, however, that an equation is a representation of a fact, and the student must be careful to avoid attempting to make his version of chemical facts agree with *guessed* equations which though apparently possible do not really represent actual reactions.

CHIEF POINTS OF CHAPTER XIX.

Atoms and Molecules.—All matter is supposed to be built up of minute particles termed *atoms*. These usually exist combined with other atoms to form *molecules*. The molecules of a compound are composed of different kinds of atoms united together. In an element the atoms are all of the same kind.

Chemical Symbols.—Chemists represent the atoms of various elements by letters which are in general the first letter, or the characteristic letters, of the English or Latin name of an element. Thus H represents the atom of hydrogen and Cl the atom of chlorine.

Avogadro's Law.—Equal volumes of all gases under similar conditions of temperature and pressure contain equal numbers of molecules.

Chemical Equations enable the chemist to represent chemical reactions by means of symbols. The formulæ of the reacting bodies are placed on the left-hand side of the equality sign, and those of the products on the right.

The **Atomic Weight** of an element is the ratio of the weight of its atom to the weight of an atom of hydrogen. Thus, when we say the atomic weight of chlorine is 35·5, we mean that the atom of chlorine is 35·5 times as heavy as the atom of hydrogen.

The **Equivalent Weight** of an element is that weight of it which can just replace 1 gram of hydrogen in a compound.

The **Valency** of an element means the number of atoms of hydrogen one atom of it is able to replace.

Atomic Weight = Equivalent Weight × Valency.

QUESTIONS ON CHAPTER XIX.

1. What do you mean by the terms "atom" and "molecule?" Carefully point out the difference between them.

2. What is Avogadro's Law? Indicate the use of this law in the determination of molecular formulæ.

3. What is meant by "atomic weight"? How does it differ from "equivalent weight"? Under what circumstances are they equal?

4. How would you endeavour to find the equivalent weight of a metal soluble in sulphuric acid?

5. By what experiments and reasoning would you show that hydrochloric acid gas should be represented by the formula HCl?

6. What is "valency"? Give examples of elements possessing different valencies.

7. What is a chemical equation? Write down the equations representing the changes of any three chemical operations you have seen, and give in words the exact meaning of the equation.

8. Write in words the *exact* meaning of the equation $2KClO_3 = 2KCl + 3O_2$. From this equation calculate the weight of oxygen obtainable from 100 grams of potassium chlorate.

9. Explain how the measurement of the density of a gas leads to the determination of its molecular weight. State in words the law upon which this determination is based.

10. Explain clearly why we represent the molecule of water by the formula H_2O.

11. "The molecule of hydrogen must contain at least two atoms." Justify this statement.

12. Sulphuretted hydrogen may be proved experimentally to contain its own volume of hydrogen. Describe the experiment, and say what it proves regarding the *molecule* of sulphuretted hydrogen.

CHAPTER XX

PHOSPHORUS AND ITS COMPOUNDS

General Characters of Phosphorus.—Phosphorus, like sulphur, exists in different allotropic forms, and these must be first examined. Ordinary phosphorus, which is always kept under water, will probably appear at first to be a yellow, or brown, opaque solid. This is not, however, its true appearance, but only that of a film which coats the exterior. If a piece be cut it will be found to consist of a waxy, translucent solid, of a pale yellow colour. It very readily inflames, contact with any warm object being sufficient to start the combustion. For this reason it should not be handled, and when not in use should be kept immersed in water, in which it is insoluble. Phosphorus glows in the dark, forming white fumes. Both glow and fumes are due to its gradual oxidation. Phosphorus itself, as well as the fumes of its oxide, is poisonous, and hence care must be taken during its use, both on this account and on account of its inflammability.

> Expt. 286.—Place a little phosphorus in an evaporating basin containing water and heat slowly. Observe the temperature at which the phosphorus is seen to melt. Allow it to cool.

You will find that the phosphorus melts at about 43° C., but remains liquid below this temperature. This is frequently the case with melted solids, but during the solidification the temperature again rises to the melting point. If the water be poured off, the phosphorus frequently catches fire, as it inflames in air at a temperature below the melting point.

> Expt. 287.—Shake up a small piece of phosphorus in carbon disulphide. Observe that it dissolves. Pour the solution on to a piece of filter paper. Observe the effect.

In this experiment the evaporation of the solvent leaves the phosphorus deposited on the filter paper in so finely divided a state that it rapidly oxidises, charring the paper, even setting it alight. By careful evaporation of the solvent crystals may be obtained.

Red Phosphorus.—This allotropic form of phosphorus is a dark, brownish-red powder. Examine it carefully and note well the differences from the ordinary variety, substituting red phosphorus, and try again the various experiments performed. It will be found to be insoluble in carbon disulphide, and to only ignite when strongly heated (240° C.). It is not luminous, neither does it oxidise when exposed to moist air. It is therefore not necessary to keep red phosphorus under water. Red phosphorus, or amorphous phosphorus, as it is also called, is, further, non-poisonous.

Phosphorus and Oxygen.—We have stated before that when phosphorus burns it forms a compound called phosphorus pentoxide. This is the product of combustion when either red or yellow phosphorus is used, and this fact may serve to prove that the two varieties are chemically identical. The composition of the oxide is represented by the formula P_2O_5, the prefix *pent* (five) indicating the number of atoms of oxygen in the molecule of the compound. The properties of this oxide have been already partially studied (p. 222). It is a white amorphous powder which dissolves very readily in water, forming an acid solution.

If left exposed to the air it absorbs moisture, being, in fact, one of the most powerful of drying agents, on account of which it is frequently used as a *dehydrator* (*i.e.*, for the purpose of removing moisture from gases or liquids). For the same reason it is of service in promoting many chemical reactions, the essential part of which is the removal of the elements of water.

Phosphoric Acid and Phosphates.—The acid solution which results on the solution of the oxide, P_2O_5, in water is known as *phosphoric acid*,[1] and its composition is represented by the formula H_3PO_4, so that we may write the equation for its production thus :—

$$P_2O_5 + 3H_2O = 2H_3PO_4.$$

We may also obtain the acid directly from phosphorus itself by means of nitric acid, which is a powerful oxidising agent (*i.e.*, a substance which readily gives up oxygen to other bodies).

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[1] This acid results if the water be hot: if cold water be used, another acid, which gradually changes to this, results.

EXPT. 288.—Place some phosphorus, either variety, in an evaporating basin, add some nitric acid, and warm gently in a draught cupboard. Notice the energetic action and the evolution of red fumes (a sign that the nitric acid is giving up oxygen). Evaporate nearly to dryness, then add a little more nitric acid and evaporate down again. You thus obtain a gelatinous product, which cools to a hard glassy mass.

The phosphoric acid so obtained is a crystalline solid, which dissolves in water. Like the other acids already examined, it may be neutralised by the addition of caustic soda, and a salt obtained. The salts obtained from phosphoric acid are known as *phosphates*. Owing to the presence of three hydrogen atoms in the molecule of phosphoric acid, we may obtain three sets of salts in which one, two, or three hydrogen atoms are replaced; thus with sodium we may obtain trisodium phosphate, Na_3PO_4; hydrogen disodium phosphate, HNa_2PO_4; and dihydrogen sodium phosphate, H_2NaPO_4. Salts like the two latter, in which all the hydrogen is not replaced are known as *acid salts;* those which, like the first, contain no acid hydrogen are termed *normal salts*.

Phosphorous Oxide.—Other oxides of phosphorus, however, exist besides the pentoxide, and of these the most important is that known as *phosphorus trioxide*, *phosphorous oxide*, or *phosphorous anhydride*, P_4O_6. This results, to a small extent, when phosphorus burns in air, the quantity increasing when the supply of air is insufficient. It is a white solid with a garlic-like odour, and it differs greatly from the pentoxide in its action on water, as it is only very slowly dissolved, forming an unstable acid, *phosphorous acid*, H_3PO_3, the salts of which are called *phosphites*.

Compounds of Phosphorus and Hydrogen.—By the action of phosphorus on caustic soda, a most remarkable product results, and the following experiments should be done to obtain and study this compound.

EXPT. 289.—Fit up apparatus in the manner indicated in Figure 134. In the retort place some pieces of phosphorus and a strong solution of caustic soda. Then completely displace the air from the apparatus by means of a current of coal gas. When this is done the retort may be heated Observe that action goes on inside and that a gas with a offensive odour, recalling that of rotten fish, is evolved. After a short time each bubble of the gas ignites directly it reaches the air, forms a bright flash and a white cloud which ascends as a cloud ring. Collect a jar of the gas, but do not remove it from the water, pass up into it a piece of litmus paper. Observe it has no action. Allow this jar to stand over water, and examine the next day. It does not now ignite when exposed to the air, but on the application of a lighted taper burns with formation of white fumes.

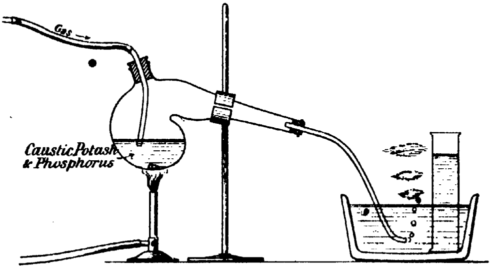

FIG. 134.—Preparation of Phosphoretted Hydrogen.

N.B.—Be very careful not to let the water rush back into the retort, otherwise it will crack and an explosion result. This is easily prevented by turning on the coal gas should the water be seen ascending from any cause.

The gas so obtained is a compound of phosphorus and hydrogen, known as **phosphine** or **phosphoretted hydrogen**, and represented by the formula PH_3. This formula may be proved by the decomposition of the gas by electric sparks, when it forms red phosphorus and hydrogen, 2 volumes of gas yielding 3 volumes of hydrogen. By the action of chlorine, also, it forms a chloride of phosphorus and hydrochloric acid, which we know contains hydrogen. The white fumes formed during its combustion consist of a mixture of water and a compound known as metaphosphoric acid, and the equation for this action may be thus written :—

$$PH_3 + 2O_2 = H_2O + \underset{\text{Meta-phosphoric Acid.}}{HPO_3}.$$

The gas is poisonous, so that care has to be taken that it is not inhaled in an appreciable quantity; the preparation is best performed in a draught cupboard or in the open air. It was seen that the gas loses its property of *spontaneous inflammability* if allowed to stand over water, and the same effect results if the gas be passed through a cooled receiver before collection. This indicates that this property is not due to the gas itself, but to some other product mixed with it, and separated from it, by either of the above methods. This other product is also a compound of phosphorus and hydrogen, but it has a composition represented by the formula P_2H_4, and is liquid at ordinary temperatures. It is, as we should expect, spontaneously inflammable, and its presence in other combustible gases, as the gaseous phosphine, hydrogen, &c., causes them, also, to ignite in air, and become themselves spontaneously inflammable.

In the retort (Expt. 289) used for the preparation of the gas there remains, together with the excess of phosphorus and caustic potash, a compound known as *sodium hypophosphite.*

Manufacture of Phosphorus.—Phosphorus is employed for various purposes, but the greatest quantity is used for the production of matches, the tips of which consist of a little wax with phosphorus and potassium chlorate. The chief source of phosphorus is the residue obtained on burning bones, which is known as *bone ash*, and consists of *phosphate of calcium*, $Ca_3(PO_4)_2$. This product is converted into phosphoric acid by treating with sulphuric acid, when the following change takes place:—

$$\begin{array}{ccccccccc} \text{Calcium phosphate} & \text{and} & \text{sulphuric acid} & \text{form} & \text{calcium sulphate} & \text{and} & \text{phosphoric acid,} \\ Ca(PO_4)_2 & + & 3H_2SO_4 & = & 3CaSO_4 & + & 2H_3PO_4. \end{array}$$

The calcium sulphate, being insoluble, is easily separated from the acid, which is then concentrated and heated with powdered coke in cast iron retorts connected with pipes dipping under water, by which means phosphorus is obtained. The reactions for this process may be thus represented:—

$$H_3PO_4 = HPO_3 + H_2O.$$
$$4HPO_3 + 12C = 12CO + 2H_2 + 4P.$$

The crude phosphorus so obtained is purified by melting under warm water, the melted phosphorus being then cast, while still under water, in the form of round sticks.

CHIEF POINTS OF CHAPTER XX.

Phosphorus, like sulphur, exists in different allotropic forms.

Ordinary phosphorus is a waxy, translucent solid of a pale yellow colour. It very readily inflames and should not, in consequence, be handled, and when not in use should be kept immersed in water. It glows in the dark, evolving white fumes. It is poisonous. It melts at about 43° C. It is soluble in carbon disulphide.

Red phosphorus is a dark brownish-red powder, insoluble in carbon disulphide, and only ignites when strongly heated. It is non-poisonous. It need not be kept in water.

Phosphorus and Oxygen.—When either ordinary or red phosphorus is burnt in air or oxygen a compound called *phosphorus pentoxide*, P_2O_5, is formed. It is a white amorphous powder which dissolves very readily in water, to form an acid solution known as *phosphoric acid*, H_3PO_4. [$P_2O_5+3H_2O=2H_3PO_4$.] Phosphoric acid may be obtained directly from phosphorus by oxidising with nitric acid.

Phosphoric acid forms salts called *phosphates*. Since phosphoric acid contains three atoms of displaceable hydrogen it is possible to obtain three different sodium salts of phosphoric acid; those salts which contain atoms of hydrogen as well as sodium are known as *acid salts*, while those in which all the hydrogen atoms have been displaced are called *normal salts*.

Phosphorous Oxide, P_4O_6, is formed to a small extent when phosphorus burns in air, the quantity increasing if the supply of air is insufficient. It is a white solid with a garlic-like odour. With water it forms phosphorous acid, H_3PO_3.

Compounds of Phosphorus and Hydrogen.—The most important of these is *phosphine* or *phosphoretted hydrogen*, PH_3. It is evolved when phosphorus is heated with a strong solution of caustic soda. It has an offensive smell of rotten fish. When prepared in the manner described in the chapter it is spontaneously inflammable, but this is due to its admixture with another compound of phosphorus and hydrogen, P_2H_4, which can be got rid of either by passing the gas through a cooled receiver or by allowing it to stand over water.

Manufacture of Phosphorus.—Phosphorus is usually prepared from the residue obtained by burning bones. This residue is a phosphate of calcium, $Ca_3(PO_4)_2$. By treating this compound with sulphuric acid it is converted into phosphoric acid. After concentration the phosphoric acid is heated with powdered coke in cast iron retorts, when phosphorus distils over and is cooled by water.

QUESTIONS ON CHAPTER XX.

1. In what respects do red and yellow phosphorus differ, and in what respects do they resemble one another? What reasons have we for believing they are chemically identical?

2. What is formed when phosphorus burns? Give an account of the appearance and properties of the product.

3. What are phosphates, and how may they be obtained?

4. How is ordinary phosphoric acid obtained? What is its effect upon caustic soda?

5. Give an account of the preparation and characteristic properties of phosphoretted hydrogen. What precautions have to be strictly observed in its preparation?

6. Describe, with diagram, an experiment to find the relative weights of phosphorus and oxygen which combine during the combustion of phosphorus.

7. When phosphorus has been kept for some time it may not have the appearance and properties which are usually ascribed to it. Why is this?

8. What oxides of phosphorus are there? How are they obtained and how do they differ?

9. For what purposes may the higher oxide of phosphorus be used in chemical work, and to what property is its use due?

10. Give a short account of the mode of production of phosphorus. For what purposes is phosphorus chiefly employed?

11. How is it that whereas only one salt is obtained by the action of caustic soda upon hydrochloric acid, three salts are obtainable from phosphoric acid?

12. Phosphoretted hydrogen as ordinarily prepared is spontaneously inflammable. It may however be so prepared that it is not spontaneously inflammable. To what is the difference due, and how may the first gas be converted into the second?

CHAPTER XXI

CARBON AND SOME OF ITS COMPOUNDS

Forms of Carbon.—Carbon is an element which is very widely distributed in nature, being present in all living matter, and in most products resulting from vital activity.

> Expt. 290.—Heat a series of organic substances, such as meat, wood, potato, egg, &c., in a crucible, and notice in all cases the production of a black residue, consisting largely of carbon. Heat more strongly, and observe that it burns away, leaving an almost colourless ash.

Carbon occurs also in many rock masses, being a constituent of all the minerals known as carbonates. Combined with oxygen as carbon dioxide, it occurs in the atmosphere and certain natural gases, or dissolved in spring waters.

In the pure state it exists in various allotropic forms. Of these the purest and the most valuable is the **Diamond.** This form of carbon is crystalline and very hard, being capable of scratching all other minerals. Its refractive index (p. 171) is very high, and on this depends its brilliancy as a gem. Diamond is proved to consist of carbon by burning it, when only *carbon dioxide* results.

Blacklead or **graphite** is another form of almost pure carbon, with properties totally different from those of the diamond. It is opaque and black, and so soft that it will mark paper. It is really a crystalline form of carbon, although good crystals are not very common. It occurs naturally in mines, chiefly in California, and was formerly largely obtained from Cumberland. Besides its use for lead pencils, it is also used as a lubricant.

Other forms of more or less pure carbon in an **uncrystallised**

3. What are phosphates, and how may they be obtained?

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6. Describe, with diagram, an experiment to find the relative weights of phosphorus and oxygen which combine during the combustion of phosphorus.

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> EXPT. 290.—Heat a series of organic substances, such as meat, wood, potato, egg, &c., in a crucible, and notice in all cases the production of a black residue, consisting largely of carbon. Heat more strongly, and observe that it burns away, leaving an almost colourless ash.

Carbon occurs also in many rock masses, being a constituent of all the minerals known as carbonates. Combined with oxygen as carbon dioxide, it occurs in the atmosphere and certain natural gases, or dissolved in spring waters.

In the pure state it exists in various allotropic forms. Of these the purest and the most valuable is the **Diamond.** This form of carbon is crystalline and very hard, being capable of scratching all other minerals. Its refractive index (p. 171) is very high, and on this depends its brilliancy as a gem. Diamond is proved to consist of carbon by burning it, when only *carbon dioxide* results.

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Other forms of more or less pure carbon in an **uncrystallised** or **amorphous** state are *coke*, and *gas carbon*, which result from the heating of coal; *lampblack*, which is the carbon deposited by oils, &c., burning in an insufficient supply of oxygen; *wood charcoal*, &c., obtained by heating wood, in closed retorts or in stacks under earth.

Charcoal has the power of absorbing many gases, and also of absorbing colouring matter, and on the latter account it is used for decolorising solutions coloured by organic matter. Coal contains large quantities of carbon, especially the harder or anthracite coals, where the quantity may reach 94 per cent., being, however, only about 65 in the brown coal or lignite. *Animal charcoal* is really a misleading term, as the quantity of carbon present is usually only about 10 or 12 per cent., the remainder being chiefly bone-ash.

Compounds of Carbon.—Carbonates.—It has been already stated that when carbon burns it forms an oxide, which is called *carbon dioxide*. This is known to be a colourless gas, with a very faint smell; it may be recognised by the turbidity it produces in lime-water. We must now study this gas further, and investigate the change it produces in lime-water; but before doing this, it will be necessary to know something concerning the latter, so that we shall first examine lime.

Lime.—Lime is a white solid, which is unchanged on heating; but which, if heated sufficiently, glows and emits a brilliant white light. It is on this account employed for the production of the limelight, where a small, hard cylinder of lime is strongly heated in an oxy-hydrogen or oxy-coal-gas flame.

> EXPT. 291.—Dissolve some lime in hydrochloric acid, and evaporate the solution to dryness. Note the formation of a white solid, which rapidly absorbs moisture from the air and liquefies. It has been previously used in some of our experiments under the name of *calcium chloride*.

Try the effects of other acids, and observe that lime yields with nitric acid a white *soluble* solid, and with sulphuric acid a white *insoluble* solid.

> EXPT. 292.—Test some wet lime with litmus papers. Observe it changes red litmus to blue.

Lime is obtained from chalk, or from limestone, the rock being strongly heated in kilns, from which, after the burning, the lime is raked out. If a piece of freshly formed lime can be obtained it should be examined.

Expt. 293.—To some fresh lime add some cold water, and observe that the mass gets very hot and swells up, still remaining a solid.

This lime is known as *quick lime*, and by this addition of water becomes in some way changed, since the further addition of water causes no further evolution of heat. The altered lime is known as *slaked lime.*

To find out the change undergone by the chalk in limestone a small portion should be heated, and the gain or loss of weight recorded. This can be done if a furnace—a muffle furnace is very convenient—is available.

Expt. 294.—Weigh a crucible and in it place some powdered chalk.[1] Heat it in a muffle furnace at a bright red heat for a few hours, allow it to cool, and again weigh. Calculate the loss or gain for 100 grams of chalk.

It will be found that the chalk loses about 44 per cent. of its weight, and it is thus proved that the lime consists of chalk from which some constituent has been expelled by heat.

Expt. 295.—To a little chalk in a test-tube add some dilute hydrochloric acid. Note brisk effervescence occurs. Fit a delivery tube to the test-tube and pass the gas evolved through lime-water, and observe the formation of a white precipitate resembling that produced by carbon dioxide. To some of the precipitate add a little dilute acid, and observe it becomes clear, with slight effervescence. Evaporate the solution left in the test-tube and see that, as in the case of the action between *lime* and hydrochloric acid, *calcium chloride* is left.

From chalk then, by the action of hydrochloric acid, we obtain carbon dioxide and the product which results from the interaction of the same acid with lime. It appears therefore probable that the chalk consists of the carbon dioxide and the lime. The next experiment will prove whether or not this is the case.

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[1] It is worth while to point out that the prepared substance used for black-board writing is often not chalk.

Expt. 296.—Into a flask A (Fig. 135) fitted with india-rubber stopper and tubes, as shown, pour some dilute hydrochloric acid, and weigh the flask with the contained acid. Weigh a small piece of tube B (i) empty (ii) with some chalk. Place this in the flask, taking care not to allow the acid to reach the inside of the tube. Replace the cork and shake the flask so as to allow the acid and chalk to react. The carbon dioxide formed escapes through the tube C in which cotton-wool is packed to retain moisture. When the action is at an end draw a current of air through the flask in order to displace the carbon dioxide still left inside. Weigh the flask again, and subtract the weight so obtained from that of the flask + weight of tube and chalk. The resulting number is evidently the weight of the carbon dioxide evolved. Calculate the loss for 100 grams of chalk.

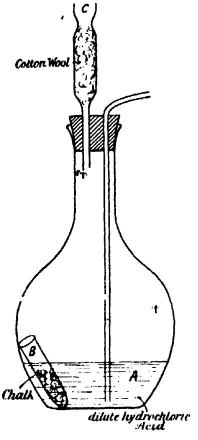

Fig. 135.—To illustrate Experiment 296.

The loss of weight of the chalk is equal to the loss of weight on heating, so that we are justified in concluding that the same constituent was evolved, and we may write :—

Chalk = lime and carbon dioxide.

By the action of the gas upon lime dissolved in water we also obtain a white product, and if sufficient of this be collected and examined it is found that by the action of acids it also gives off the carbon dioxide, and that it forms lime on heating. It is therefore chalk, and the formation is merely the reverse of what was written above, *i.e.*, lime and carbon dioxide form chalk, so that the turbidity of the lime-water is due to the formation of chalk.

Compounds which, like chalk, evolve carbon dioxide when acted upon by an acid, are known as **carbonates**, and many carbonates exist, all possessing many characteristics in common. Many of these on heating also give off carbon dioxide, the esidue being known to be an oxide of a metal, so that they onsist of carbon dioxide and a metallic oxide.

EXPT. 297.—Heat copper carbonate, and observe the production of black oxide of copper.

This leads us to the idea that lime is also the oxide of a netal, and this view is now known to be correct, the metal eing named *calcium*. Lime therefore is *calcium oxide*, CaO, nd chalk *calcium carbonate*, $CaCO_3$. The action of carbon lioxide upon lime-water may be written

$$CaO + CO_2 = CaCO_3.$$

nd the decomposition of chalk by heat is given by the equation :

$$CaCO_3 = CaO + CO_2.$$

Carbon Dioxide.—We are now in a position to prepare larger uantities of pure carbon dioxide, and to study its properties nore fully.

EXPT. 298.—In a flask fitted with cork thistle funnel, and delivery tube (Fig. 136), place some chalk or marble (also a form of calcium carbonate). Pour in some dilute hydrochloric acid, and collect the gas over water as in the case of the preparation of hydrogen. Collect three jars.

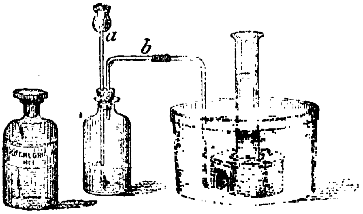

FIG. 136.—The Preparation of Carbon Dioxide.

Pass the gas through a solution of lime-water. Observe the formation of the white precipitate which afterward disappears. Boil the solution so obtained, and notice the re-precipitation.

In the first jar place a lighted taper. Observe the eff

Shake up with water, and pour the water into some blue litmus solution.

Pour the gas from one jar into another, as shown in the diagram (Fig. 137), and test both jars by a lighted taper. It will be seen that the lower jar contains the gas.

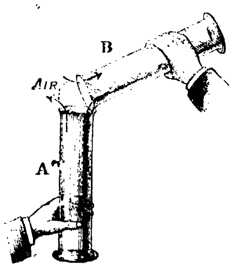

FIG. 137.—To illustrate the High Density of Carbon Dioxide.

Properties of Carbon Dioxide.—It is thus seen that carbon dioxide is a heavy gas which does not burn nor support combustion. It is slightly soluble in water, the solution acting as a weak acid and turning blue litmus to a port wine colour, different from the red produced by strong acids.

The solution of the gas may in fact be regarded as containing an acid, carbonic acid, of composition H_2CO_3 (thus, $CO_2 + H_2O = H_2CO_3$), from which the carbonates are derived.

The formation of the white precipitate with lime-water has been shown to be due to the formation of calcium carbonate. Its disappearance is usually explained by the statement that the calcium carbonate is soluble in a solution of carbon dioxide, although insoluble in pure water, and is hence re-precipitated when the gas is boiled away. The true explanation most probably is that really a new compound is formed from the calcium carbonate and the carbonic acid, viz., $CaH_2C_2O_6$, hydrogen calcium carbonate, which is soluble in water, and which decomposes when heated, again forming chalk.

Uses of Carbon Dioxide.—The fact that carbon dioxide is a non-supporter of combustion is made use of in many forms of chemical fire extinguishers, which are generally solutions which on heating evolve carbon dioxide in which the burning cannot take place and is therefore extinguished.

The solubility of carbon dioxide in water is increased by pressure, and the sparkling nature of the various aërated waters is due to the carbon dioxide with which they have been charged at high pressures, and which escapes when the pressure is reduced to that of the atmosphere by the opening of the bottle. It is also produced during fermentation, the "rising" of bread being due to the escape of carbon dioxide generated by the fermentation of the saccharine matters formed from the starch under the influence of the yeast.

By cold and pressure carbon dioxide may be liquefied and also solidified, forming a soft white substance, which when mixed with ether forms a powerful freezing mixture, the temperature sinking to about -100° C.

Occurrence of Carbon Dioxide.—Carbon dioxide has been already stated to exist in the atmosphere, and to be produced by the oxidation of animal and vegetable tissues, while further, under the influence of sunlight it is reconverted by the green parts of plants into its constituents, of which the carbon is used by the plant in the formation of tissue. It is also found in many natural gaseous emanations, and is frequently present to a large extent in the gases of caves and underground passages, where, owing to its high density, it tends to accumulate, if formed by fermentation or other natural processes. In expired air carbon dioxide is present to the extent of about 4·7 per cent. Although such air is not again respirable, this is partly due to the diminution of the oxygen, and it is doubtful whether carbon dioxide has any direct poisonous effect. The proportion of carbon dioxide may be increased to even 20 per cent. without immediate serious effects if only the quantity of oxygen be simultaneously increased.

Proof of formula CO_2.—The density of carbon dioxide should be found in the manner indicated for hydrochloric acid gas, and will be found to be 22. As, further, the combustion of carbon in oxygen leads to no change of volume, the formula CO_2 for the gas, with the atomic weight of carbon = 12, is indicated.

Hardness of Water.—EXPT. 299.—To a solution of calcium carbonate or chalk, in water containing carbon dioxide, add some soap solution and shake. Observe that a great quantity of soap must be added before a lather is obtained. Boil the solution first, and again add the soap. Observe a lather is readily formed.

The solution is hence a very *hard water*, but becomes soft again by boiling. The **temporary hardness** may then be due to chalk dissolved in water by the aid of carbon dioxide, and the reason for the loss of hardness on boiling is evident, since the chalk is then precipitated and no longer exists dissolved in the water. Temporary hardness is in fact generally due to chalk, but another closely allied compound, magnesium carbonate, gives similar results. We may say, however, that it is due to the *carbonates* of *calcium* or *magnesium* dissolved in water containing carbon dioxide ; and the white solid obtained in Expt. 250 was probably one of these compounds.

Expt. 300.—Shake a solution of calcium chloride with soap solution. Observe that this is also very *hard*, but that the hardness is not removed by boiling. The water is *permanently* hard.

Permanent hardness is also generally due to a calcium salt (or magnesium salt) soluble in water itself, and hence not precipitated by boiling ; calcium sulphate or magnesium sulphate is the salt most commonly found in permanently hard water. In either case, *i.e.*, temporary or permanent hardness, the calcium (or magnesium) unites with a portion of the soap and hence destroys it, forming a precipitate which renders the water turbid. Until sufficient soap is used to unite with the whole of the calcium present no lather is obtainable.

Lower Oxide of Carbon.—If a coke or charcoal fire be observed, flames of a peculiar blue colour are usually seen playing over the top. As, however, the only substances concerned in the combustion are carbon and oxygen, and since carbon itself does not burn with a blue flame, while carbon dioxide is quite incombustible, it appears probable that there must be some other compound of carbon and oxygen which is produced and which burns with these blue flames. Further, it seems probable that the oxide is a lower oxide, *i.e.*, contains less oxygen, as if a compound with more oxygen existed we should expect the carbon dioxide to burn and form this oxide. We may therefore first endeavour to obtain this oxide by abstracting some of the oxygen from carbon dioxide.

Expt. 301.—In a hard glass tube AB, such as was used in Expt. 234, place some iron filings. Pass over these filings a current of carbon dioxide, dried in the usual manner. Heat the filings, and let the end B of the tube dip under a strong solution of caustic potash, which absorbs any unchanged carbon dioxide, a tube containing the same liquid being inverted over the end (Fig. 138.) Observe that the bubbles of gas which come from the end of the tube are *almost* completely absorbed but that a gas slowly collects and fills the tube. Take out the tube and apply a light to the gas. Observe that it burns with a blue flame.

We have, therefore, obtained the gas expected, and if the tube be examined oxide of iron is seen to be present. On burning, the gas again forms carbon dioxide, and its composition may be represented by CO, the name **carbon monoxide**[1] being given to it. We may then state :—

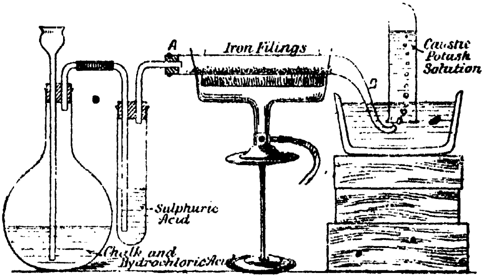

FIG. 138.—Preparation of Carbon Monoxide.

Carbon dioxide and iron yield carbon monoxide and iron oxide, or

$$4CO_2 + 3Fe = 4CO + Fe_3O_4.$$

The burning of carbon monoxide may be also expressed by the following equation:—

$$2CO + O_2 = 2CO_2.$$

Changes in a Coke Fire.—The formation of this gas in coke fires may now be readily explained. By the burning of the carbon in the lower parts of the fire, where the air enters and sufficient oxygen is present, carbon dioxide results, but when this passes over the heated charcoal in the upper parts where insufficient oxygen is present, it parts with some of its oxygen, forming carbon monoxide. Thus,

$$CO_2 + C = 2CO.$$

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[1] It is sometimes spoken of as carbonic oxide, but the name used above is preferable.

When it again meets sufficient oxygen at the top of the fire, this monoxide burns to the dioxide. Carbon monoxide may in fact be prepared by passing the dioxide over heated charcoal instead of heated iron, but the temperature required is higher than that necessary in the experiment described.

The gas so obtained is colourless and very poisonous. It acts as a direct poison, since it replaces the oxygen from its compound with the hæmoglobin of the blood. For this reason the ventilation of rooms heated by coke stoves, and the complete removal of the products of combustion should be always well attended to.

Compounds of Carbon with Hydrogen.—Carbon also forms compounds with hydrogen, direct union occurring when the electric arc is produced between carbon poles in an atmosphere of hydrogen. The compound resulting is known as **Acetylene**, and has the composition represented by C_2H_2. It burns with a highly luminous flame, and is present to a small extent in the products of the incomplete combustion of coal gas. It is now prepared largely by the action of water on *calcium carbide*, and its cleanliness and ease of preparation will probably lead to its wider use as an illuminant.

Ethylene, another compound of carbon and hydrogen, is a product of the destructive distillation of coal, wood, &c. It is hence present as a constituent of coal gas. It is readily obtained by heating a mixture of alcohol and sulphuric acid. Like acetylene, it burns with a bright, but less luminous flame. At low temperatures it may be liquefied, and by allowing the liquid ethylene to rapidly evaporate (page 155), the temperature is so greatly reduced that air may be liquefied. The composition is represented by C_2H_4, and the following equation may be given for its preparation:—

$$\underset{\text{Alcohol}}{C_2H_5OH} = \underset{\text{Ethylene.}}{C_2H_4} + H_2O.$$

In this only the final products are considered, intermediate compounds are, however, really first formed between the alcohol and the acid.

Marsh Gas.—Marsh gas, or methane, has the composition CH_4, and occurs naturally, being a product of the decomposition of vegetable matter. It is found rising to the surface in swamps, and thus obtains its name marsh gas. It is also fre quently found in coal mines, and is known as *fire-damp*. It is an inflammable gas, and hence its presence is a source of danger in collieries, and many disastrous explosions have been caused by its ignition.[1] Water and carbon dioxide (the after-damp) are produced by its combustion, and the effects of the products in causing suffocation are often more fatal than those of the explosion itself, while the presence of small quantities of carbon monoxide adds greatly to the loss of life.

In the laboratory, methane is usually prepared by heating a mixture of sodium acetate and caustic soda. It is colourless, odourless, insoluble in water, and burns with feebly luminous flame.

Ordinary coal gas consists of a mixture of various hydrocarbons, chiefly marsh gas, together with hydrogen. By its combustion carbon dioxide and water are formed. If sufficient air be introduced into the flame it has been seen (Chap. XV.) that this combustion proceeds without the separation of any carbon. If air be not introduced into the gas some of the hydrocarbons become decomposed—the hydrogen burning, to form water, and the carbon being separated. This carbon is raised to incandescence and afterwards burned, when there is sufficient air.

CHIEF POINTS OF CHAPTER XXI.

Carbon is present in all living matter. When organic substances are moderately heated, a black residue largely composed of carbon is left. If the temperature is raised, the carbon burns away and an almost colourless ash is left.

Carbon exists in several allotropic forms. Two of these, *diamond* and *graphite*, are crystalline. Non-crystalline or *amorphous* carbon is known in varying degrees of purity as *coke*, *gas carbon*, *wood charcoal* and *animal charcoal*.

Lime, which is used in making *lime-water*, is a white solid which is unchanged on heating. But when heated intensely it glows and emits a brilliant white light. It dissolves in hydrochloric acid to form *calcium chloride*. Wet lime turns a red litmus paper blue. Lime is obtained from chalk and limestone by strongly heating them. When freshly made it is called *quick lime* which combines with water, evolving much heat and becoming converted into *slaked lime*.

When **Chalk or Limestone is heated** it loses about 44 per cent. of its weight. This is due to the loss of carbon dioxide. The same gas is evolved when either of the substances is acted upon by hydrochloric acid. By ascertaining the weight per cent. of carbon dioxide which is given off by the chalk, we can prove that chalk is really a compound of lime and carbon dioxide.

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[1] Finely divided coal dust has also been shown to be capable of giving rise to explosions, and it is possible that many regarded as due to fire-damp were in reality due to coal dust, which appears to produce more violent explosions than the gas.

Carbon Dioxide is a heavy gas which does not burn nor support combustion. It is slightly soluble in water, the solution acting as a weak acid which turns blue litmus to a port wine colour. This solution may be regarded as *carbonic acid*, H_2CO_3. The solubility of carbon dioxide in water is increased by pressure. By cold and pressure it may be liquefied and also solidified to a soft, white substance, which when mixed with ether forms a powerful freezing mixture.

Carbon dioxide is always found in the atmosphere, being produced by the oxidation of animal and vegetable tissues. This carbon dioxide is decomposed by the green colouring matter of plants in the presence of sunlight. It is very common in caverns in limestone regions. Its formula is CO_2.

Explanation of Hardness of Natural Waters.—*Temporary hardness* is due either to chalk (calcium carbonate) or to magnesium carbonate which is dissolved in the water. The solution of these compounds is possible because the carbon dioxide in the water first converts them into *soluble hydrocarbonates.* When the water is boiled these compounds are converted into insoluble carbonates.

Permanent hardness is generally due to calcium or magnesium sulphate which is dissolved in the water. They cannot be got rid of by boiling because they are both soluble in water.

Carbon Monoxide contains less oxygen than the dioxide. It can be prepared from the dioxide by abstracting some oxygen. This can be effected by passing it over heated iron :—

$$4CO_2 + 3Fe = 4CO + Fe_3O_4.$$

Carbon monoxide burns with a blue flame, forming carbon dioxide again : $2CO + O_2 = 2CO_2$. It is a colourless gas which is very poisonous.

Compounds of Carbon with Hydrogen.—*Acetylene*, C_2H_2, is formed by the direct union of carbon and hydrogen. It burns with a highly luminous flame, and is present to a small extent in the products of the incomplete combustion of coal gas. It is now largely prepared by the action of water on calcium carbide.

Ethylene, C_2H_4, is a product of the destructive distillation of coal, wood, &c., and is hence found in coal gas. It can be readily prepared by heating a mixture of alcohol and sulphuric acid. It burns with a luminous flame less bright than that of acetylene. Liquid ethylene has been obtained.

Marsh Gas, CH_4, is also known as *methane.* Since it is one of the products of the decomposition of vegetable matter it is often found rising to the surface of swamps, hence the name marsh gas. It occurs in coal mines as *fire-damp.* It is generally prepared in the laboratory by heating a mixture of sodium acetate and caustic soda, when it is seen to be a colourless and odourless gas which burns with a feebly luminous flame.

Coal Gas consists of a mixture of hydrogen, carbon monoxide, marsh gas, and other hydrocarbons. In its ordinary combustion carbon is at first separated and is a cause of the luminosity of the flame.

QUESTIONS ON CHAPTER XXI.

1. Give an account of the different modifications of carbon. How may it be proved that they all really consist of the same substance?

2. What compounds of carbon and oxygen exist? How may they be obtained, and how may they be mutually converted into one another?

3. When carbon dioxide is passed into lime water, the latter becomes first turbid and afterwards clear. Explain the chemical nature of the causes of these changes.

4. Give a brief account of the part played by carbon dioxide in the economy of nature.

5. Explain the cause of the blue flames which are frequently seen playing at the top of a coke or charcoal fire.

6. If a gas containing a mixture of the two oxides of carbon were given you, how could you determine the quantity of each present?

7. Describe three experiments to show the high density of carbon dioxide, its effect on combustion, and its action on lime-water.

8. By what means would you prove the presence of carbon dioxide in atmospheric air and expired air?

9. If carbon dioxide be passed over heated iron filings what other product is formed? How would you collect it, and in what particulars does it differ from carbon dioxide?

10. What is the chemical nature of chalk and lime? How could you obtain lime from chalk and *vice versa*?

11. Describe an experiment to determine the quantity of carbon dioxide which is evolved by the action of acid upon one gram of chalk.

12. What is marsh gas? Why is it so called? What are its properties?

13. What is (*a*) "fire-damp," (*b*) "after-damp"? What are the dangers due to the presence of much fire-damp in mines?

14. By what tests would you ascertain whether effervescence caused by pouring sulphuric acid on a powder was due to the escape of carbon dioxide, sulphur dioxide, hydrochloric acid gas, or hydrogen?

15. What is the meaning of hard and soft water? How would you test to see whether a water is hard or soft?

16. Why are some waters "permanently" hard, and others only temporarily so? Explain clearly what is meant by these expressions.

17. Rain water is soft, but spring water is generally hard. By what process has the soft rain water, after falling, become hard before issuing as a spring?

18. What does coal gas consist of? What products are formed by its combustion, and how would you endeavour to experimentally prove your answer?

19. Describe the apparatus you would use for the production and collection of carbon dioxide gas; name the materials required, and describe the properties of the gas.

20. Charcoal burnt in air or oxygen in a bottle containing lime-water produces a white precipitate in the lime water. The white

precipitate, if collected and mixed with hydrochloric acid, dissolves with effervescence. What experiments would you make in order to compare the gas thus obtained with that obtainable from the breath ?

21. Three bottles filled with colourless gas are placed before you containing (*a*) hydrogen, (*b*) nitrogen, and (*c*) carbon dioxide. Describe experiments you would make in order to distinguish them from one another, and from other gases.

22. Describe and explain the changes which take place when (*a*) limestone is burnt in a kiln, (*b*) water is added to some freshly burnt lime.

23. How is lime distinguished from limestone as to composition and properties ?

CHAPTER XXII

ACIDS, BASES, AND SALTS

Acids and Salts.—It has already been shown that from *hydrochloric acid* there can be obtained common salt and other allied compounds, called *chlorides*, while from *sulphuric acid* we get a series of compounds called *sulphates*. It will be necessary to consider a little more fully the formation of these compounds.

Besides the two acids mentioned above, a third is very frequently employed in the laboratory, viz., *nitric acid*, HNO_3. Examine this acid, and notice that it gives off pungent, choking fumes, especially on heating, being then coloured yellow. By the action of copper, red fumes and a blue liquid result; this forms a good means of recognising the acid. If the blue liquid be evaporated and allowed to crystallise, a salt which closely resembles the copper sulphate is obtained, which is called *copper nitrate*, consisting of nitric acid in which the hydrogen has been replaced by copper.

We may therefore consider **acids as compounds containing hydrogen which is capable of being replaced by a metal with the formation of salts.** They have usually a sour taste and change the colour of blue litmus to red. **Salts** may be evidently defined as **compounds resulting from acids by the replacement of their hydrogen by a metal.**

Preparation of Salts.—We can frequently prepare salts by the action of the acid directly on the metal, and numerous examples of such action have been studied. In the first place hydrogen is liberated, and this hydrogen may either pass off as a gas or may react again with the acid to form various products. An example of the first case has been seen in the action of sulphuric acid on zinc, and of the second case in the action of nitric acid on copper. If, however, an oxide of the metal be used instead of the metal itself, such secondary actions are avoided, as the following experiments show.

EXPT. 302.—Burn some magnesium, and collect the white solid formed. Observe that it is a white powder insoluble in water, which glows when strongly heated, but does not undergo any chemical change.

To some of this oxide add some diluted sulphuric acid, but not quite enough to dissolve all the oxide. Observe that the oxide dissolves *without* the evolution of gas. Filter, partially evaporate, and allow it to crystallise.

The white crystals so obtained consist of *magnesium sulphate.* They possess a bitter saline taste, and may be easily proved to contain water of crystallisation (p. 239). They are commonly known as *Epsom Salts*, and are used as an aperient. In this case, then, no hydrogen was evolved, since the oxygen of the oxide was capable of combining with it to form water ; thus :—

$$\underset{\text{Magnesium Oxide.}}{MgO} + H_2SO_4 = \underset{\text{Magnesium Sulphate.}}{MgSO_4} + H_2O.$$

Similar results may be obtained if nitric or hydrochloric acid be used in place of sulphuric acid, with the formation of *magnesium nitrate*, $Mg(NO_3)_2$, and *magnesium chloride*, $MgCl_2$:—

$$MgO + 2HNO_3 = Mg(NO_3)_2 + H_2O.$$
$$MgO + 2HCl = MgCl_2 + H_2O.$$

Calcium Salts.—Other metallic oxides can in a similar manner react with acids, with the production of salts and water.

EXPT. 303.—Examine lime (oxide of calcium). Note its effect on litmus, its slight solubility, action of heat, &c. (page 286), and, as above, endeavour to prepare the calcium sulphate, calcium chloride, and calcium nitrate.

You will find that crystals do not result, but that *calcium chloride* and *calcium nitrate* are both white solids which readily absorb moisture from the air. Owing to this the former is, as we have seen, largely used for drying gases, &c. The *sulphate*, however, does not dissolve in water (at least, it dissolves only to a very small extent, viz., about 2 parts per 1,000), and hence the acid may at first be thought to have no action, but the action is readily proved to take place by the loss of its characteristic properties by the acid. Owing to its insolubility it is also formed as a precipitate, when sulphuric acid is added to a fairly strong solution of calcium chloride or calcium nitrate. It is largely used technically, being the substance known as *gypsum* or *plaster of Paris*, which possesses the property of setting to a hard mass when mixed with water. The formation of these salts may be represented by the equations:—

$$CaO + 2HNO_3 = Ca(NO_3)_2 + H_2O.$$
$$CaO + 2HCl = CaCl_2 + H_2O.$$
$$CaO + H_2SO_4 = CaSO_4 + H_2O.$$

Reactions of Hydroxides with Acids.—It has been shown that caustic soda and caustic potash are also capable of forming salts. These compounds are, however, not oxides, but are known as *hydroxides*, *i.e.*, compounds resulting from the displacement of only half the hydrogen in water, or compounds containing one or more OH groups. Caustic soda is thus *sodium hydroxide*, NaOH, and caustic potash is *potassium hydroxide*,[1] KOH. The latter compound may be prepared by placing the metal potassium on water. In this case the action is far more energetic than is the case with sodium, and the metal appears to burn on the surface of the water with a violet flame. In reality it is the liberated hydrogen which burns, the violet colour being due to vapours of potassium.

EXPT. 304.—Place a small piece of potassium on water, and afterwards evaporate the liquid to dryness. Examine the solid produced, and compare it with caustic soda.

Caustic soda and caustic potash will be found to resemble one another closely. They are both white solids, with a peculiar soapy feel, both turn red litmus blue, and both absorb moisture and carbon dioxide very readily from the air.

EXPT. 305.—Make solutions of caustic potash and caustic soda, and neutralise both by (1) sulphuric acid, (2) nitric acid, (3) hydrochloric acid. If care be taken the solution may be made almost perfectly neutral, *i.e.*, neither acid nor alkaline. Evaporate the solutions so obtained and allow them to crystallise.

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[1] These compounds are also called sodium hydrate and potassium hydrate, but the term hydroxide is preferable.

The following salts are thus obtained: *potassium sulphate*, *potassium nitrate*, *potassium chloride*, *sodium sulphate*, *sodium nitrate*, and *sodium chloride*. All are white solids, two of them being important commercial products, viz., potassium nitrate, which is known as *nitre* or *saltpetre*, and sodium chloride, which is *common salt*.

Quantitative Nature of the Reactions.—It will be also necessary to examine these reactions quantitatively.

Expt. 306.—Weigh out *about* 40 grams of pure caustic soda and about 50 grams of pure caustic potash. (To do this a weighed and corked flask may be used, since these compounds rapidly absorb moisture from the air.) Dissolve each in distilled water and make up the solution to exactly 1 litre, using a measuring flask for this purpose. Make up also dilute solutions of sulphuric, nitric, and hydrochloric acids; about 1 volume of the strong laboratory acid to 20 volumes of water.

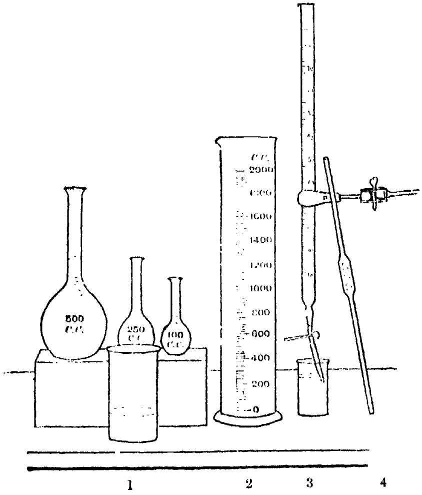

Fig. 139.—Volumetric Apparatus.
1. Measuring Flasks. 2. Graduated Jar. 3. Burette. 4. Pipette.

Fill a *burette* with the caustic soda solution, and place 25 c.c. of the sulphuric acid solution in a flask. This is done by *carefully* sucking up the acid into a *pipette* until the liquid passes the 25 c.c. mark, then speedily closing the top of the pipette with the finger, and allowing, by slight motion of the finger, the liquid to run out until it is exactly at the mark. To this acid add a drop of litmus, whereby the solution is coloured a bright red. Now run in the caustic soda solution from the burette until the acid is exactly neutralised, the point of neutralisation being known by the change of colour from red to blue, which occurs when the smallest excess of the caustic soda is added. Note carefully the number of cubic centimetres of caustic soda solution required to neutralise the acid. Do the experiment a second time to verify your result, and then perform a similar pair of experiments with the other two acids, and with the caustic potash and the three acids. Your results may be entered thus :—

25 c.c. of	c.c. of caustic soda required.	c.c. of caustic potash required.
Sulphuric acid.......	15·4	17·3
Nitric acid	18·2	20·4
Hydrochloric acid ...	20·0	22·4

As, however, the strength of the solutions of caustic soda and caustic potash are known, we may instead of the number of c.c. of solution write the number of grams of the dissolved compound. Thus suppose that 40 grams of caustic soda and 50 grams of caustic potash were used, then 1 c.c. of the solution contains respectively ·04 and ·05 grams, as that the above table becomes

25 c.c. of	Grams of caustic soda required.	Grams of caustic potash required.
Sulphuric acid.......	·616	·865
Nitric acid	·728	1·020
Hydrochloric acid ...	·800	1·120

Hence we find in the above suppositious experiment that ·86 gram of caustic potash can neutralise as much *sulphuric acid* as ·616 gram of caustic soda, &c. Calculate from your own experiments what quantity would neutralise as much of each acid as 1 gram of caustic soda. The result should be found to be the same for each acid, viz., that 1·4 grams of caustic potash neutralises the same quantity of *any* acid, as does 1 gram of caustic soda.

> EXPT. 307.—Weigh out also some concentrated sulphuric acid, which may be regarded as containing 98 per cent. of the pure acid, and make up a solution so that 1 litre contains about 50 grams of sulphuric acid. Again perform the neutralisation experiment, and, knowing the *weight* of actual acid present in the 25 c.c. and the *weight* of caustic soda present in the solution of caustic soda used, calculate the weight of acid which neutralises 1 gram of caustic soda. You should find that 1·225 grams of acid are necessary.

As the strength of the laboratory solution of the acids varies, you cannot weigh these, but if a solution of known strength is given, you could, in a similar manner, prove that the 1 gram of caustic soda or 1·4 grams of caustic potash can neutralise 1·575 grams of nitric acid and ·91 gram of hydrochloric acid. If we now calculate the results for the molecular weight of caustic soda we find 40 grams of caustic soda are equivalent to 56 grams of caustic potash and neutralise 49 grams of sulphuric acid, 63 grams of nitric acid or 36·4 grams of hydrochloric acid.

We may represent the changes occurring in these reactions by the following equations, and the reader should write in words the exact meaning of these (and all other) equations, and should see that the quantitative relations expressed by the equations are those actually found.

$$2KOH + H_2SO_4 = K_2SO_4 + 2H_2O.$$
$$2NaOH + H_2SO_4 = Na_2SO_4 + 2H_2O.$$
$$KOH + HNO_3 = KNO_3 + H_2O.$$
$$NaOH + HNO_3 = NaNO_3 + H_2O.$$
$$KOH + HCl = KCl + H_2O.$$
$$NaOH + HCl = NaCl + H_2O.$$

We thus see that hydroxides as well as oxides may react with acids to form salts and water. These compounds are both included under the name of **bases.** We may hence write

Base + Acid = Salt + Water,

and may define a base as **the oxide or hydroxide of a metal, capable of neutralising an acid with the formation of a salt and water.**

Alkalies.—The *soluble* bases, the student has probably observed, are characterised by *alkaline* properties, *i.e.*, are known as *alkalies.*

Insoluble hydroxides can be prepared by the addition of an *alkali* to the solution of a metallic salt. Thus, if caustic soda solution be added to, say magnesium sulphate, a double decomposition occurs with the formation of the insoluble *magnesium hydroxide*, $Mg(OH)_2$, as a white precipitate and sodium sulphate which remains in solution.

$$2NaOH + MgSO_4 = Na_2SO_4 + Mg(OH)_2.$$

On strongly heating this hydroxide, the oxide itself is obtained,

$$Mg(OH)_2 = MgO + H_2O.$$

The slaking of lime is due to the conversion of the calcium oxide into calcium hydroxide,

$$CaO + H_2O = Ca(OH)_2$$

Acid Salts and Basic Salts.—If the quantity of base is insufficient to completely neutralise the whole of the acid, we frequently obtain **acid salts** (p. 280), as acid sodium sulphate, $NaHSO_4$, thus :—

$$NaOH + H_2SO_4 = NaHSO_4 + H_2O.$$

Acid salts are also called *hydrogen salts.*

In other salts, on the contrary, a greater quantity of the base is present than is required for the neutralisation of the acid. Such salts are termed **basic salts.**

Ammonia.—Certain other compounds, however, besides the metallic oxides, may act as bases, thus we may obtain salts from **ammonia,** which, by testing with red litmus, is seen to be an **alkali.**

Expt. 308.—Prepare salts from the three acids, using the ammonia liquid of the laboratory in place of caustic soda, &c.

There are thus obtained salts which are known as **ammonium salts**, viz., ammonium sulphate, &c.

It will be therefore well to examine the "ammonia liquid," which will be already known to have a very sharp pungent smell, and to colour red litmus blue.

EXPT. 309.—Place some ammonia liquid in a flask and boil, allowing the gas that is evolved to pass over quicklime or solid caustic potash in order to dry it. Collect as shown in the accompanying diagram (Fig. 140). Observe the gas has the same smell as the liquid and similarly acts on red litmus.

In one jar place a lighted taper.

Place a second jar in water, and note the rapid absorption of the gas and rise of the water in the jar.

Dip a rod in hydrochloric acid and hold over a jar of the gas. Observe the formation of white fumes.

Shake up a jar of the gas with a little water. Examine the solution. See that it is identical with the "ammonia liquid," from which the gas was obtained, and that it loses its odour on boiling, the gas being evolved.

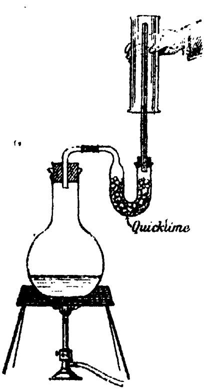

FIG. 140.—Preparation of Ammonium.

Composition of Ammonia.—We thus find the "ammonia liquid" is merely a solution of a gas, viz., ammonia, which is very soluble in water, which turns red litmus blue, and which does not support combustion, and is apparently not combustible. In an atmosphere of oxygen however the gas readily burns, forming nitrogen and water. It hence contains nitrogen and hydrogen, and may in fact be proved to have the composition represented by NH_3. To prove this, a long tube is filled with chlorine gas and by means of a funnel and the stopcock some strong liquid ammonia is run into the gas (Fig. 141). Combination ensues between the chlorine and the hydrogen of the ammonia with the formation of white fumes, and frequently a flash of light. The stopcock is then opened under water when the water [illegible] and fills two-thirds of the volume of the tube, the gas remaining being nitrogen. As chlorine combines with an equal volume of hydrogen the volume of nitrogen in the ammonia is only one-third that of the hydrogen, that is, the gas is NH_3, or N_2H_6, &c. The density of the gas (8·5), however, indicates that the simplest formula must be taken.

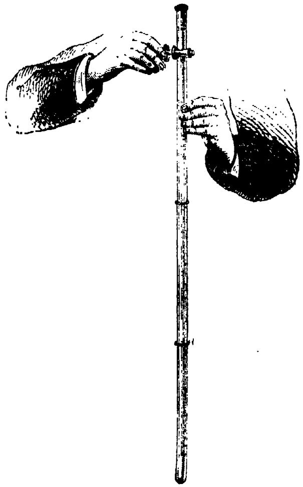

FIG. 141.—Tube for Determining the Composition of Ammonia.

The white fumes formed in the above experiment and by the action of the gas on hydrochloric acid are fumes of ammonium chloride, NH_4Cl, and its production may be represented by the equation

$$NH_3 + HCl = NH_4Cl.$$

The solution of ammonia has been already shown to behave in a manner closely analogous to caustic soda, and if added to a solution of a magnesium salt, it also precipitates magnesium hydroxide. The solution therefore behaves like a soluble hydroxide and we may regard it as such, viz., as NH_4OH, which results from the solution of the gas in water:

$$NH_3 + H_2O = NH_4OH.$$

Ammonium.—The formation of salts from ammonia is then completely analogous to the formation of salts from caustic soda, and the equation may be rewritten with the substitution of NH_4 for the Na, thus:

$$NH_4OH + HCl = NH_4Cl + H_2O.$$

[illegible] this group of elements NH_4, which we may regard as existing in these salts, the name *ammonium* is given, but it should be remembered that it only exists in its salts and has no independent existence.

Expt. 310.—Heat an ammonium salt with caustic soda or lime, or merely well mix the two in a mortar with the addition of a little water. Observe by the smell that ammonia is produced.

The gas ammonia, then, is very readily driven off from ammonium salts by the action of caustic soda, &c., this action is readily understood from the equation

$$NH_4Cl + NaOH = NaCl + NH_4OH.$$

The latter product is an aqueous solution of ammonia, so the equation may also be written thus:

$$NH_4Cl + NaOH = NaCl + NH_3 + H_2O.$$

Chief Points of Chapter XXII.

Acids are compounds containing hydrogen which is capable of bein replaced by a metal with the formation of salts, *e.g.*, hydrochloric sulphuric and nitric acids.

Salts are compounds which result from acids by the replacemen of their hydrogen by a metal.

Preparation of Salts.—Salts may be prepared either by the direc action of an acid on a metal, or by the action of an acid on an oxide In the former case hydrogen is sometimes liberated as such, while i other cases it reacts with some of the acid, forming secondary product When an oxide is used water is formed in place of hydrogen.

Examples of Salts.—The following salts among others are prepare by means of the experiments described, copper nitrate, the sulphat of magnesium, which is also known as Epsom salts, the chloride an nitrate of magnesium, and similarly the sulphate, chloride, a nitrate of calcium.

Reactions of Hydroxides with Acids.—Salts are also formed by t action of acids upon the hydroxides, such as, sodium hydroxid NaOH, or, as it is also called, caustic soda. Since oxides a hydroxides neutralise acids to form salts they are both included the class of compounds called **Bases**.

Bases are compounds, known as oxides or hydroxides of meta which are capable of neutralising acids with the formation of sa and water.

Alkalis are those bases which are soluble in water. They a characterised by an alkaline reaction. Sodium hydroxide, potassi hydroxide, &c., are examples.

Acid Salts and Basic Salts.—*Acid salts* are those which are formed when the quantity of base is insufficient to completely neutralise the whole of the acid on which it acts. Acid sodium sulphate, $NaHSO_4$, and those sodium phosphates containing hydrogen mentioned on p. 305 are examples. *Basic salts* are those which result by the action of a greater quantity of base than is required for the complete neutralisation of the acid on which it acts.

Ammonia.—When the gas, ammonia, is dissolved in water, the solution which results acts as a base of the kind described as an alkali. The salts which it forms are known as ammonium salts, *e.g.*, ammonium sulphate. When the solution of ammonia in water is heated the dissolved gas is again driven off and can be dried and collected.

The composition of ammonia is shown by the symbol NH_3. That this is so can be proved by the experiment described in the chapter.

Ammonium.—The salts which a solution of ammonia in water forms with acids are exactly analogous to those similarly formed from sodium and potassium hydroxide. This leads to the name *ammonium hydroxide* for the solution of ammonia in water. The group *ammonium*, NH_4, has no independent existence, but is only known in combination.

QUESTIONS ON CHAPTER XXII.

1. Describe the terms acid, base, salt, and indicate how they are mutually related.

2. What are normal salts, acid salts, basic salts? Give examples of an acid salt.

3. How would you obtain potassium nitrate and sodium chloride from their respective acids and bases?

4. What is ammonia? Give an account of its preparation and more characteristic properties.

5. In what respect does a solution of ammonia resemble a solution of caustic soda or caustic potash?

6. What do you mean by "ammonium" and "ammonia"?

7. Draw the apparatus, and describe how you would prepare ammonia gas from the ordinary solution of ammonia.

8. What is the difference between the action of an acid, say hydrochloric acid, upon iron and oxide of iron?

9. How would you obtain crystallised sal-ammoniac from a solution of ammonia? How would you also obtain the solution of ammonia from this product?

10. What is meant by an alkali? Compare the chemical behaviour of an alkali with that of an acid. Describe the appearance of any two alkalis.

11. You are required to prove that with regard to its power in neutralising acids fifty-six grams of caustic potash are equivalent to forty grams of caustic soda. Describe exactly how you would perform the experiment.

12. In the above question does the particular acid employed affect the result? Describe experiments to prove the truth of your answer.

13. If you were given some nitric acid and some caustic potash, describe *exactly* how you would proceed in order to obtain large crystals of nitre.

14. What are hydroxides? Give examples of metallic hydroxides. What products are formed when these hydroxides react with acids?

CHAPTER XXIII

CALCIUM CARBONATE AND SILICA

Occurrence of Carbonates.—We have already seen that chalk consists of calcium carbonate. The production of lime from this source, together with its reconversion into calcium carbonate, have been already studied in Chapter XXI. It occurs naturally in several distinct forms, and many rock masses are largely or entirely built up of this compound. When pure it occurs crystalline as the minerals *calcite* and *aragonite*, which differ only in their crystalline forms and physical properties. Calcite is also known as *Iceland spar*, *calc spar*, and by other local names. It is generally quite transparent and somewhat resembles quartz, from which it can be distinguished by its inferior hardness. It is easily scratched by a knife, while quartz is unaffected. It is what is called a *doubly-refracting* substance. If a clear crystal of Iceland spar be placed upon the page of a book and the print viewed through it, two images of each word will be seen. Calcium carbonate also occurs more or less pure in the earth's crust in a great variety of forms, such as *chalk*, *limestone*, *stalactite*, *stalagmite*, *travertine*, &c., some of which have been formed by purely chemical means, others by the aid of living organisms.

Those formed by chemical means owe their formation to the power, possessed by water containing dissolved carbon dioxide, of dissolving calcium carbonate, which is again precipitated when carbon dioxide escapes.

Travertine or Calcareous Tufa is precipitated by springs which lose their dissolved carbon dioxide, which is necessary for the solution of the calcium carbonate, as they flow onwards. The carbonate, being insoluble in water alone, is deposited as soon as the carbon dioxide escapes.

Stalactites and Stalagmites.—The streams traversing limestone districts become saturated with carbonate of lime. In their course they often trickle through crevices in the roofs of caverns which have been formed in the limestone by the same solvent power of this particular water. The drop of water, which is thereby exposed on the roof, is subjected to evaporation, and the escape of the carbon dioxide and loss of water cause a slight deposition of carbonate on the roof, which is continuously added to by a constant succession of drops, until eventually beautiful pendants of calcium carbonate are formed, called ***stalactites***, sometimes coloured by the presence of traces of iron oxide, and often having a wild profusion of forms (Fig. 142). Further evaporation of the water takes place on the floor of the cavern, giving rise to layers of the same chemical compound, called ***stalagmites***.

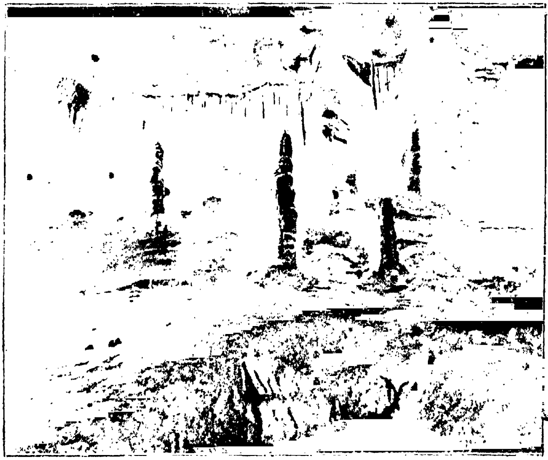

FIG. 142.—Stalactites and Stalagmites in Clapham Cave, Lancaster. From a photograph by Mr. George Fowler.

Chalk, limestone, and coral are composed chiefly of calcium carbonate, and consist almost entirely of the remains of animal organisms.

Marble, also, is a form of calcium carbonate. It is really limestone, which, by the effect of great pressure and heat in the earth's crust, has become much harder and crystalline.

Silica.—More widely spread throughout the earth's crust even than calcium carbonate is the compound silica, which consists of the oxide of an element termed *silicon*, and which is represented by the formula SiO_2. This oxide enters into the formation of by far the greater number of minerals and rocks, being combined with metallic oxides in the form of *silicates*; indeed, in the free state or thus combined, silica forms more than half the weight of the earth's crust. In the pure state it occurs both crystalline and amorphous. Two crystalline varieties are known, one, *tridymite*, is unimportant, while the other, *quartz*, is a frequently-occurring and highly-interesting mineral. If the quartz is quite clear and transparent, it is known as *rock-crystal*, and is the *Brazilian pebble* from which lenses are made for eyeglasses, &c. Sometimes the oxide of a heavy metal is present, colouring the quartz.

Many *sands* are made up entirely of grains of quartz which have become more or less rounded by continual rubbing against one another in water. If the sand becomes compacted by the introduction of a cement and by the action of great pressure a *sandstone* is formed.

Amorphous Silica.—This is found in the form of (1) chalcedony and its varieties, (2) jasper and its varieties, (3) opal.

Chalcedony is known having all sorts of tints. It is often regarded as a mixture of quartz and opal; it is familiar as the well-known red stone used in signet-rings and called *carnelian*. *Agate* is a variegated chalcedony, composed of different coloured bands. *Flint* is generally of a black or dark gray colour, and is found in nodules or bands in the chalk formations of Surrey, Kent, &c.

Jasper is an opaque, impure form of silica, of a red, brown, or yellow colour.

Opal.—This species of naturally-occurring silica always contains water in varying amounts. It is supposed to contain some quartz as well as amorphous silica. It is often used in jewellery;

one variety shows a well-developed iridescence, and is called *precious opal.*

Rocks formed of Silica.—Besides these minerals some rock-masses are composed of silica. Of these, as in the case of calcium carbonate, some have been deposited chemically, others by plant life or animal organisms.

Sinter is the rock resulting from the deposition of silica dissolved in the water of hot springs and geysers. It is, however, probable that, even in this case, minute plants of the *Algæ* group assist in the formation.

The organically formed rocks of silica may be divided into :—

(1) *Those formed by Plants.*—These plants also belong to the Algæ group and are known as the *diatoms.* They are generally microscopic in size, but have existed in sufficient numbers to form by the accumulation of their hard remains, beds of considerabl thickness. They are made of the silica which the living diatom secreted from the water, either fresh or salt, in which it lived At Richmond, in Virginia, beds, of forty feet thick, and consisting entirely of diatoms, occur. *Diatomaceous earths* and *Tripol powder* have been made in this way.

(2) *Those formed by Animals.*—These silica-secreting animal belong to the class of simplest animal structure known.; the are called *Radiolaria.* Their remains build up the *Radiolaria earths* which occur in various places, notably Barbados.

Chemically Prepared Silica.—When chemically prepared silica, or *silicon dioxide,* is obtained as a white powder, which insoluble both in water and in most acids. If fused with causti soda, however, it yields a glassy mass, which is soluble in wate and is known as *water glass.* Chemically it consists of *sodiu silicate,* Na_2SiO_3 ; in this, as in the natural silicate, the silica pla the part of an acid oxide.

As a rule the silicates are insoluble and their composition very complicated. Little is known concerning their exact chemic nature. Amongst the commoner silicates may be mentioned *cla* a silicate of aluminium ; *felspar,* a silicate of aluminium ar potassium ; *serpentine, talc, French chalk,* and *meerschaum* silicates of magnesium ; and *asbestos,* a silicate of calcium ar magnesium.

CHIEF POINTS OF CHAPTER XXIII.

Calcium Carbonate is very largely found in the earth's crust. It occurs pure and crystalline in *calcite* and *aragonite* and in a less pure condition in *chalk*, *limestone*, *stalactites*, *stalagmites*, *travertine*, etc.

Chalk, Limestone and Coral are composed chiefly of calcium carbonate, and consist almost entirely of the remains of animal organisms.

Marble also is a form of calcium carbonate. It is really limestone which by the effects of great pressure and heat in the earth's crust has become much harder and crystalline.

Silica, SiO_2, is even more widely found than calcium carbonate. It constitutes, either free or combined with bases to form silicates, about one-half the weight of the earth's crust.

In the free state it occurs both crystalline and amorphous. Two crystalline varieties are known, one, *tridymite*, is unimportant. The other, **quartz**, is very abundant.

Many *sands* are made up entirely of grains of quartz.

Amorphous Silica is found in three forms (1) chalcedony, (2) jasper, (3) opal. Carnelian, agate, and flint are common kinds of chalcedony.

Rocks formed of Silica.—*Sinter* results from the deposition of the silica dissolved in the water of hot springs and geysers.

Diatomaceous Earths and Tripoli Powder are rocks composed of silica which was extracted from the water in which they lived by minute *plants* known as *diatoms*.

Radiolarian Earths are rocks composed of silica which was extracted from the water in which they lived by minute *animals* known as *Radiolaria*.

Chemically prepared Silica is a white powder which is insoluble both in water and in most acids. Fused with caustic soda it forms *sodium silicate*, which is soluble in water and is known as *water-glass*.

QUESTIONS.

1. What is silica? How does it occur in nature?
2. How do flint and quartz differ from and resemble one another?
3. Give a short account of the naturally occurring forms of silica.
4. What is calcite? Describe its appearance and give the formula representing its chemical composition.
5. Many rocks are deposited through the agency of living organisms. Mention some and briefly indicate how they are thus deposited.
6. What are stalactites and stalagmites? Of what do they consist and how are they formed?
7. Describe three naturally occurring forms of calcium carbonate.
8. What effect has chalk upon water which flows through or over it?

9. Mention three minerals consisting of silica, and state by what characteristic feature you would recognise them.

10. How would you attempt to prove that limestone, marble, and calcite were essentially similar compounds? [The student is advised to consider well what experiments he has previously carried out with calcium carbonate.]

11. Limestone is dissolved in hydrochloric acid and the solution evaporated. Briefly describe the appearance and properties of the substance which you would expect to be thus obtained.

ANSWERS

[In every case π has been taken as equal to $\frac{22}{7}$].

CHAPTER III.

8. 23·04 ft. per sec.
11. 180 ft. per sec.
12. 49 ft. per sec. per sec.
13. 20 kilom. per hour.
14. 40·5 ft.
15. $5\frac{13}{15}$ ft. per sec. per sec.
16. 18·25 . . . ft. per sec.
17. 900 ft. ; 11 secs.
18. $3\frac{1}{3}$ ft. per sec. per sec ; 6 secs.
19. 36·33 . . . secs.
20. $\frac{1}{4}$ mile.
21. $\frac{1}{2}$ hr. ; 20 miles, 15 miles.
22. 6 secs.

CHAPTER IV.

10. 1612 cub. cm., 1·612 litres, 1612 grams.
11. 1712 cub. cm., Mass of 1 c. c. in grams $= \frac{25,000}{1500}$ 16·6.

CHAPTER V.

3. 7 : 8 ; 7 : 8.
4. 33 : 20.
8. 2 : 7.
13. 720 ft.
14. 64 ft. per sec.
15. 225 ft.
16. 144 ft.
17. (*a*) 32. ft. per sec., (*b*) 64 ft. per sec., (*c*) 96 ft. per sec.
18. 3 secs ; 96 ft. per sec.
21. 156 lbs. wt.

22. 17 units ; 7 units ; 13 units.
23. 12·2 units ; 125° nearly.
24. 65 lbs.
25. 3 lbs., 4 lbs.
26. 3P + 4Q, 2P + 3Q.
27. 5 lbs.
28. 63 P.
29. 53° nearly.
30. 21·8 units ; 13·2 units.
31. $10\sqrt{5}$ lbs.
32. 15·7 lbs.
33. 10 lbs.
34. 67·26 lbs.
35. $16\sqrt{3}$ lbs., $8\sqrt{3}$ lbs.
37. 5 : 1.
38. The second by $\frac{1}{2}$ sec.
39. $5\frac{1}{3}$ ft. per sec. per sec. ; 16 ft. per sec.
40. 75 : 7.
41. 5 : 2.
42. $2\frac{1}{2}$ lbs. wt.
43. 7 poundals.
44. $6\frac{1}{4}$ lbs. wt. ; $\frac{5}{56}$ ft. per sec. per sec.
45. equal.
46. 1 : 8.
47. 96 ft.
48. (*a*) $3\frac{1}{3}$ ft. per sec. per sec. ; (*b*) 48 lbs.
49. $10\frac{1}{8}$ lbs. wt.
50. (*a*) 180 lbs. wt. ; (*b*) 160 lbs. wt.
51. $10\frac{5}{8}$ lbs. wt.
52. 95.
53. 9856 : 3.
54. 25 : 11.
55. 3 ft. per sec.
56. $9\frac{49}{81}$ ft. per sec.
57. 8 ft. per sec.
58. 36 lbs.
59. (*a*) 1600 ; (*b*) 160.
60. 321 : 322.
61. 2 secs.
62. $3\frac{1}{3}$ lbs. wt.
63. 10 lbs.
64. Magnitude of resultant 5.
66. (*a*) $6 + 5 = 11$; (*b*) $6 - 5 = 1$.

CHAPTER VII.

12. 36·42 c.c. ; 7·55.
13. 4·63 oz. nearly.
14. ·068 nearly.
15. $\frac{75}{1024}$; $\frac{81}{16}$.
16. 9800 cub. in.

17. 11,200 cub. in.
18. 4·2
19. 128·135.

CHAPTER IX.

6. 12,000 ; $\frac{5}{12}$ths.
7. 200.
8. 16 hrs.
9. 118,800 ft. lbs.
10. $942\frac{6}{7}$ ft. lbs.
11. 24,000 ft. lbs.
12. 250 ft. poundals or $\frac{125}{16}$ ft. lbs.
13. 300 ft. lbs ; 300 ft. lbs.
14. 2250 ft. lbs. ; $120\sqrt{2}$ ft. per sec.

CHAPTER X.

19. Temperature of the mixture would have been less.
20. Water will melt more ice than the lead.
21. (*a*) 60 degrees, (*b*) 55 degrees.

CHAPTER XVI.

7. 15 vols.

ERRATUM.

P. 42, Question 8. For "five miles an hour," read, "five *times* an hour."

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